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ABSTRACT

This syllabus presents material to be covered in chemistry for pupils of varying skills and abilities in New York public high schools. An outline of topics provides an overview of the principles, concepts, and applications included within the course and tested in the New York State Regents Examination. Twelve units of study are described. Material within each unit is organized into three subheadings: topics, understandings and fundamental concepts, and supplementary information. Unit 12 outlines topics representing minimum requirements for skills and activities in laboratory experiences. (CS)

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TRIAL DRAFT

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REGENTS CHEMISTRY SYLLABUS

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Albany, New York

FOREWORD

Fifteen area meetings were held throughout New York State to discuss the contents of the Regents Chemistry Syllabus during the spring of 1977. A committee of teachers was convened later in the year to collate the information gathered from the field. The members of this committee were: Frank Eichler, New Hyde Park Memorial High School; Beulah Patton Durr, William Nottingham High School; Richard Quest, Oneonta High School; Ronald N. Ulrich, Avon Central School; Martin L. Franklin, Franklin Academy; I. Fred Socolof, Grover Cleveland High School (NYC); Anthony A. Galitsis, Sleepy Hollow High School; Harry A. Kranepool, Bishop Loughlin High School; Peter E. Demmin, Amherst Central School; and Lys Keneally, John Adams High School.

Based upon the information gathered, a Regents Chemistry Advisory Committee met in June 1978 to help set guidelines for the revision of the Regents Chemistry Syllabus. The intent of this committee was to establish a course of study having greater flexibility than the present course without lengthening the time requirement. The Regents Chemistry Advisory Committee consisted of: David W. Crane, Greece-Arcadia High School; Beulah Patton Durr, William Nottingham High School; Anthony A. Galitsis, Sleepy Hollow High School; Larry C. Grimm, Clarkstown Senior High School North; Lys Keneally, John Adams High School; Chirakkal V. Krishnan, East Islip High School; Joseph J. Oberkrieser, Orchard Park High School; I. Fred Socolof, Grover Cleveland High School (NYC); Edith M. Stillwaggon, Division Avenue Senior High School; and Marlene Williams, New Hartford High School

It was virtually impossible to get substanial agreement on what items must be deleted from the present syllabus in order to include some additional items requested by the field. Therefore, the committee suggested a format which would essentially maintain the present content of the syllabus and at the same time incorporate the suggestions from the field. This was accomplished without altering the minimum requirements of the course.

During the summer of 1978 the writing team of C. V. Krishnan of East Islip High School and J. J. Oberkrieser of Orchard Park High School prepared a Development Draft for Regents Chemistry Syllabus according to the guidelines established by the Regents Advisory Committee. Copies were mailed to all schools in October with a request for reactions to the draft and suggestions for its improvement.

L. Keneally and B. Durr of the Regents Chemistry Advisory Committee collated the responses from the school districts, categorizing them for examination by the committee. The Developmental Draft for the Regents Chemistry Syllabus had been received favorably by a majority of teachers in the field. This made it necessary for the committee, meeting in June 1979, to recommend only a few further guidelines for the preparation of the Trial Draft for the Regents Chemistry Syllabus.

Each and every response from the various school districts was also read by the writing team. Many suggestions made by teachers were included in the present Trial Draft of the syllabus. The end result has been a reorganization of the present syllabus which provides some options for teachers, resulting in a flexibility that was not present before. Some new topics have been added. Wherever possible, content has been clarified and updated according to modern concepts.

With the large degree of variability in different situations encountered in New York State, this syllabus allows the teacher the decision to study selected units in greater depth and still have time to complete the minimum requirements for Regents examination in chemistry. Teachers are not limited by the scope of this draft of the syllabus which lists the minimal requirements for the purpose of testing. They may add enrichment beyond the required areas.

The Developmental Draft and the Trial Draft were prepared in the Bureau of General Education Curriculum Development under the direction of Robert F. Zimmerman, Associate in Secondary Curriculum. John V. Favitta, Associate in Science Education, supervised the development of both drafts.

INTRODUCTION

OVERVIEW

This course of study presents a modern view of chemistry suitable for pupils with a wide range of skills and abilities. The outline of topics provides the unifying principles of chemistry. The principles included in the outline are basic to the understanding of our environment.

The technological impact of chemistry should not be the only driving motivation for the study of chemistry. Students should also be made aware of the total effect of the application of chemical principles on our lives. Teachers are encouraged to use those applications which they deem important and which have local or intrinsic significance.

PREREQUISITES

In order to be able to understand and use the principles of chemistry introduced in this outline, students should be familiar with the use of standard notation of numbers, significant figures, metric system of units, heat units, dimensional analysis (inclusion of units in mathematical computations), and an understanding of direct and inverse relationships.

Students enrolling in the Regents chemistry course must have completed Ninth Year Mathematics, Course I - (Algebra). They also should have completed or, at least, be currently enrolled in Tenth Year Mathematics. While very little of the content of Mathematics 10 is used directly in chemistry, the experiences in setting up and solving problems, and the analytical thinking developed in this course are most useful in chemistry.

TEACHING SEQUENCE

Teachers are at liberty to use any sequence of topics outlined on the next page and such teaching techniques as are appropriate for their students.

Teachers should not hesitate to introduce briefly any of the subject matter as the occasion demands, even though it will be treated in depth later in the course. The teaching should be an interlocking process so that the students get both a foretaste of what is to come and a review of the material previously covered.

LABORATORY

Laboratory work should be designed to encourage students to search for relationships. To do this, some laboratory exercises must be quantitative in nature. It is suggested that individual and class results of laboratory data be analyzed from time to time so that students can develop concepts of precision and accuracy. The type of experiment sometimes referred to as "cook-book chemistry" offers little toward the development of understanding.

A chemistry laboratory exercise is defined here as the laboratory work done by the student during one school period. In addition, a satisfactory written report of this work is required. The minimum laboratory requirement can be met by performing 30 individual exercises requiring 30 laboratory periods. The minimum requirement may also be met by 30 periods involving a smaller number of experiments which may require two or more periods for completion. This is not intended to permit the student to spread what is ordinarily a one-period exercise over two or more periods in order to earn additional time credit for this exercise.

The unit 12* on Laboratory Activities presented in this syllabus is not intended to include all of the Regents chemistry laboratory requirements outlined above. Rather, the topics outlined in this unit represent the minimum requirements for testing in the Regents examination if this unit is elected by the students.

TIME REQUIREMENT

The minimum time required for this course is six 40-minute periods per week, although seven periods are recommended. The minimum length of a period shall be 40 minutes, exclusive of the time used in changing the classes or teachers. This time allotment should include at least one double period for laboratory work each week.

STATE DIPLOMA CREDIT

This course may be used as one unit of the Group II major science sequence or for Group III credit as an elective toward a State diploma.

SYLLABUS SCOPE AND CONTENT

The material in this syllabus is organized under three headings: Topics, Understandings and Fundamental Concepts, and Supplementary Information. Materials presented under all the headings will be included in the Regents examination. All the material organized under the headings, Activities, Minimum Requirements and Supplementary Information in Unit 12*, Laboratory Activities, will also be subject to the Regents examination.

Units 1-9 are required of all students. In addition a minimum of two Units from Units 5*, 6*, 8*, 9*, 10*, 11* and 12* are required of all students.

Unit 1.	Matter and Energy	Unit 7.	Acids and Bases
Unit 2.	Atomic Structure	Unit 8.	Redox and Electrochemistry
Unit 3.	Bonding	Unit 8*.	Additional Materials in Redox and Electrochemistry
Unit 4.	Periodic Table	Unit 9.	Organic Chemistry
Unit 5.	Mathematics of Chemistry	Unit 9*.	Additional Materials in Organic Chemistry
Unit 5*.	Additional Materials in Mathematics of Chemistry	Unit 10*.	Applications of Chemical Principles
Unit 6.	Kinetics and Equilibrium	Unit 11*.	Nuclear Chemistry
Unit 6*.	Additional Materials in Kinetics and Equilibrium	Unit 12*.	Laboratory Activities

EXPERIMENTAL EXAMINATION

Part I of the experimental examination will contain 65 objective questions for which 65 credits will be allowed. The pupil is to answer all questions on Part I. Part I of the experimental examination will be based on Units 1-9. Those Units with an asterisk (*) are not included on Part I of the examination.

Part II of the experimental examination will consist of twelve groups of objective questions based on each of the Units of the entire syllabus. Each group will be worth five credits. The pupil is to choose seven of the twelve groups and answer all questions in each group chosen. A total of 35 credits will be allowed for Part II.

THIS TENTATIVE FORMAT FOR THE EXAMINATION WILL BE USED ONLY BY THE TRIAL SCHOOLS FOR THE JUNE 1980 EXAMINATION.

The student in the experimental course will have the choice of receiving either the raw score on the experimental examination or a grade determined by adding the raw score on the examination to the student's class average for the year and dividing by two. The higher of the two marks may be selected for use on the permanent record as the student's official Regents grade in chemistry.

TOPICAL OUTLINE

Units 1-9 are required of all students. A minimum of two units with an asterisk (*) are required of all students. Numbers in parenthesis refer to page location.

UNIT 1 - MATTER AND ENERGY (1)

I. Definition of Chemistry (1)

II. Matter (1)

A. Substances (1)
1. Elements
2. Compounds

B. Mixtures (1)

III. Energy (1)

A. Forms of energy (1)
B. Energy and chemical change (1)
1. Exothermic reaction
2. Endothermic reaction

C. Measurement of energy (2)
1. Calorie
2. Thermometry
a. Fixed points on a thermometer

IV. Phases of Matter (2)

A. Gases (3)
1. Boyle's law
2. Charles' law
3. Standard temperature and pressure (STP)
4. Partial pressures
5. Kinetic theory
6. Deviations from the gas laws
7. Avogadro's hypothesis

B. Liquids (5)
1. Vapor pressure
2. Boiling point
3. Heat of vaporization

C. Solids (6)
1. Crystals
2. Melting point
3. Heat of fusion
4. Sublimation

UNIT 2 - ATOMIC STRUCTURE (7)

I. Atoms (7)

A. Introduction to atomic structure (7)

B. Important subatomic particles (7)
1. Electrons
2. Nucleons
a. Protons
b. Neutrons

C. Structure of atoms (7)
1. "Empty space" concept
2. Nucleus
a. Atomic number
b. Isotopes
c. Mass number
d. Atomic mass (weight)
3. Electrons

D. Atomic structure models (8)
1. Principal energy levels
2. Quanta
3. Spectral lines

E. Orbital model of the atom (9)
1. Energy levels
a. Principal quantum numbers
b. Sublevels
c. Orbitals
2. Electron configurations
3. Valence electrons

F. Ionization energy (11)

II. Natural Radioactivity (11)

A. Differences in emanations (12)
1. Alpha decay
2. Beta decay
3. Gamma radiation

B. Separating emanations (12)
C. Detection of radioactivity (13)
D. Half-life (13)

UNIT 3 - BONDING (14)

I. The Nature of Chemical Bonding (14)

A. Energy stored in bonds (14)
B. Energy changes in bonding (14)
C. Bonding and stability (14)
D. Electronegativity (14)

II. Bonds Between Atoms (14)

A. Ionic (15)

B. Covalent (15)
1. Nonpolar covalent
2. Polar covalent
3. Coordinate covalent
4. Molecular substances
5. Network solids
C. Metallic (17)

III. Molecular Attraction (17)

A. Dipoles (17)
B. Hydrogen bonding (17)
C. Van der Waals forces (17)
D. Molecule-ion attraction (18)

IV. Directional Nature of Covalent Bonds (18)

V. Chemical Formula (18)

A. Symbol (18)
B. Formula (19)
1. Molecular
2. Empirical

VI. Naming and Writing Formulas of Chemical Compounds (19)

VII. Chemical Equations (20)

UNIT 4 - PERIODIC TABLE (21)

This material need not be taught as a unit, but may be incorporated in various places in the syllabus at the discretion of the teacher.

I. Development of Periodic Table (21)

II. Properties of Elements in the Periodic Table (21)

A. Covalent atomic radius (21)
B. Ionic radius (22)
C. Metals (22)
D. Nonmetals (22)
E. Metalloids (22)

III. Chemistry of a Group (Family) (23)

A. Groups IA and IIA (23)
B. Groups VA and VIA (24)
C. Group VIIA (26)
D. Group 0 (26)
E. Transition elements (27)

TOPICAL OUTLINE

IV. Chemistry of a Period (27)

UNIT 5 - MATHEMATICS OF CHEMISTRY (28)

This material need not be taught as a unit, but may be incorporated in various places in the syllabus at the discretion of the teacher.

I. Mole Interpretation (28)

II. Use of the Mole Concept (28)

A. Gram atomic mass (gram-atom) (28)
B. Gram molecular mass (28)
C. Molar volume of a gas (28)

III. Stoichiometry (28)

A. Problems involving formulas (29)
1. Percent composition
2. Empirical formula
B. Problems involving equations (29)
1. Mass problems
2. Mass-volume problems
3. Volume problems

IV. Solutions (30)

A. Methods of indicating concentrations (30)
1. Molarity

UNIT 5* - ADDITIONAL MATERIALS IN MATHEMATICS OF CHEMISTRY (32)

Teachers who elect this unit 5* may wish to teach materials outlined here in the appropriate places in the syllabus.

I. The Mole - Additional Problems (32)

II. Formula from Percent Composition (33)

III. Molecular Mass from Gas Density (34)

IV. Effect of Solute on Solvent (34)

A. Boiling point elevation (34)
B. Freezing point depression (34)
C. Abnormal behavior of electrolytes (35)

V. Calorimetry (35)

A. Heat of vaporization - problems (35)
B. Heat of fusion - problems (35)

VI. Combined Gas Laws (35)

VII. Graham's Law (35)

UNIT 6 - KINETICS AND EQUILIBRIUM (36)

I. Kinetics (36)

A. Role of energy in reaction rates (36)
1. Activation energy
2. Heat (enthalpy) of reaction
3. Potential energy diagram
B. Factors affecting rate of reaction (37)
1. Nature of the reactants
2. Concentration
3. Temperature
4. Surface areas
5. Catalysts

II. Equilibrium (39)

A. Phase equilibrium (39)
B. Solution equilibrium (39)
1. Gases in liquids
2. Solids in liquids
3. Solubility
C. Chemical equilibrium (40)
1. LeChatelier's principle
a. Effect of concentration
b. Effect of pressure
c. Effect of temperature
d. Effect of catalyst
2. Law of chemical equilibrium

III. Spontaneous Reactions (43)

A. Energy changes (43)
B. Entropy changes (43)

UNIT 6* - ADDITIONAL MATERIALS IN KINETICS AND EQUILIBRIUM (44)

Teachers who elect this unit 6* may wish to teach materials outlined here in the appropriate places in the syllabus.

I. Free Energy Change (44)

II. Predicting Spontaneous Reactions (44)

III. Solubility Product Constant (K_{sp}) (45)

A. Common ion effect (45)

TOPICAL OUTLINE

UNIT 7 - ACIDS AND BASES (46)

I. Electrolytes (46)

II. Acids and Bases (46)

A. Acids (46)
1. Arrhenius' theory
2. Brönsted-Lowry theory
B. Bases (48)
1. Arrhenius' theory
2. Brönsted-Lowry theory
C. Amphoteric (amphiprotic) substances (48)

III. Acid-Base Reactions (48)

A. Neutralization (48)
1. Acid-base titration
2. Salts
B. Conjugate acid-base pair (49)

IV. Ionization Constant (50)

A. K_w (50)
B. pH (51)

UNIT 8 - REDOX AND ELECTROCHEMISTRY (52)

I. Redox (Oxidation-Reduction) (52)

A. Oxidation (52)
B. Reduction (52)
C. Oxidation number (52)
D. Redox reactions (54)

II. Electrochemistry (54)

A. Half-reactions (54)
B. Half-cells (54)
C. Chemical cells (Electrochemical cells) (54)
D. Electrolytic cells (54)

III. Balancing Simple Redox Equations (55)

UNIT 8* - ADDITIONAL MATERIALS IN REDOX AND ELECTROCHEMISTRY (56)

Teachers who elect this unit 8* may wish to teach materials outlined here in the appropriate places in the syllabus.

I. Standard Electrode Potentials (56)

A. Half-cell potential (56)
B. Use of standard electrode potentials (56)
C. Equilibrium (58)

II. Chemical Cells - Calculations (58)

III. Electrolytic Cell Reactions (59)

IV. Electrodes (59)

A. Cathode (59)
B. Anode (59)

V. Electroplating (59)

VI. Balancing Redox Equations (60)

UNIT 9 - ORGANIC CHEMISTRY (61)

I. Definitions (61)

II. Characteristics of Organic Compounds (61)

A. Bonding (61)
B. Structural formulas (62)
C. Isomers (62)
D. Saturated and unsaturated compounds (62)

III. Homologous Series of Hydrocarbons (62)

A. Alkanes (63)
B. Alkenes (63)
C. Alkynes (63)
D. Benzene series (63)

IV. Other Organic Compounds (64)

A. Alcohols (65)
1. Primary alcohols
B. Organic acids (65)

V. Organic Reactions (66)

A. Substitution (66)
B. Addition (66)
C. Fermentation (67)
D. Esterification (67)
E. Saponification (67)
F. Oxidation (67)
G. Polymerization (67)

TOPICAL OUTLINE

UNIT 9* - ADDITIONAL MATERIALS IN ORGANIC CHEMISTRY (68)

Teachers who elect this unit 9* may wish to teach materials outlined here in the appropriate places in the syllabus.

I. Alcohols (68)

A. Monohydroxy alcohols (68)
1. Primary alcohols
2. Secondary alcohols
3. Tertiary alcohols
B. Dihydroxy alcohols (68)
C. Trihydroxy alcohols (69)

II. Aldehydes (69)

III. Ketones (70)

IV. Ethers (70)

V. Polymers (70)

A. Condensation (70)
B. Addition (71)

UNIT 10* - APPLICATIONS OF CHEMICAL PRINCIPLES (72)

Teachers who elect this unit 10* may wish to teach materials outlined here in the appropriate places in the syllabus.

I. Chemical Theory and Industry (72)

II. Industrial Applications (72)

A. Equilibrium and reaction rates (72)
1. Haber process
2. Contact process
B. Redox (73)
1. Reduction of metals
2. Corrosion
3. Batteries
a. Lead-acid battery
b. Nickel-cadmium battery
C. Petroleum (75)
1. Fractional distillation
2. Cracking

UNIT 11* - NUCLEAR CHEMISTRY (77)

Teachers who elect this unit 11* may wish to teach materials outlined here in the appropriate places in the syllabus.

I. Artificial Radioactivity (77)

A. Artificial transmutation (77)
1. Accelerators

II. Nuclear Energy (77)

A. Fission reaction (77)
1. Fuels
2. Moderators
3. Control rods
4. Coolants
5. Shielding

B. Fusion reaction (78)
1. Fuels
2. High energy requirement

C. Radioactive wastes (79)

D. Uses of radioisotopes (79)
1. Based on chemical reactivity
2. Based on radioactivity
3. Based on half-life

UNIT 12* - LABORATORY ACTIVITIES (81)

This unit is not intended to include all of the Regents Laboratory requirements. See Introduction to Syllabus, p. ii.

I. Measurement (81)

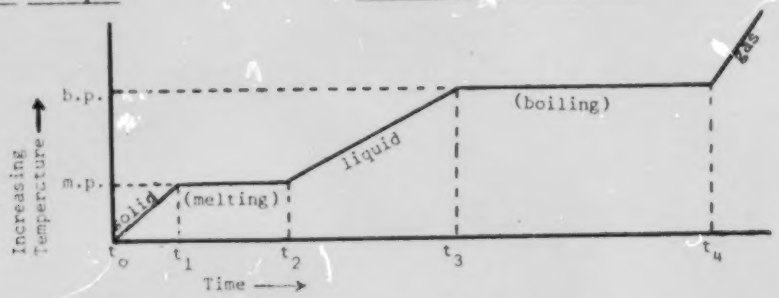
II. Laboratory Skills (82)

III. Laboratory Activities (83)

IV. Laboratory Reports (84)

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Definition of Chemistry	Chemistry is the study of the composition, structure and properties of matter, the changes which matter undergoes, and the energy accompanying these changes.	
II. Matter		
A. Substances	A substance is any variety of matter all specimens of which have identical properties and composition.	All samples of a particular substance have the same heat of vaporization, melting point, boiling point, and other properties related to composition which can be used for identification.
	A substance is homogeneous.	*For the purpose of Regents examinations, the slight differences in composition of a substance due to differences in isotopic masses will not be considered.*
1. Elements	An element is a substance which cannot be decomposed by chemical change.	All samples of an element are composed of atoms with the same atomic number.
2. Compounds	A compound is a substance which can be decomposed by a chemical change. A compound is composed of two or more different elements chemically combined in a definite ratio.	All samples of a compound have identical composition. Binary compounds are compounds consisting of only two elements.
B. Mixtures	A mixture consists of two or more distinct substances differing in properties and composition. The composition of a mixture can be varied.	Mixtures may be homogeneous (example, solutions or mixtures of gases) or heterogeneous (example, a mixture of iron and sulfur).
III. Energy		
A. Forms of energy	Heat, light, and electricity are forms of energy.	
	Energy may be converted from one form to another but is never destroyed in a change.	Examples should include reactions involving a variety of forms of energy.
B. Energy and chemical change	Energy is either given off or absorbed in most chemical changes.	Energy absorbed by the reactants may be stored in the bonds which hold atoms together. This energy may be liberated when stronger bonds are formed. This concept is developed in Units 3 and 6.
		Although it is sometimes convenient to distinguish between physical changes and chemical changes, the distinction is not clear cut and to most scientists is of little importance.

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Exothermic reaction	Energy is released in an exothermic reaction.	The role of energy in reactions is developed in Unit 6 and Unit 10*.
2. Endothermic reaction	Energy is absorbed in an endothermic reaction.	
C. Measurement of energy	Because energy in various forms may be converted to heat, the chemist uses heat units (calories or kilocalories) to measure the energies involved in chemical reactions.	
1. Calorie	One calorie is the amount of heat required to raise the temperature of one gram of water one Celsius degree.	*For scientific uniformity the term "Celsius" will be used.*
	One kilocalorie is equivalent to 1000 calories.	*For the purpose of Regents examinations this definition of calorie will be used.*
2. Thermometry	Temperatures are indicators of the direction in which heat will flow. Heat flows spontaneously from a body at higher temperature to a body at a lower temperature.	The temperature of a body is a measure of the average kinetic energy of its particles. At the same temperature, the average kinetic energy of the particles of all bodies is the same.
a. Fixed points on a thermometer	The fixed points on a thermometer are the ice-water equilibrium temperature at 1 atmosphere pressure, and the steam-water equilibrium temperature at 1 atmosphere pressure.	
	On the Celsius scale the ice-water equilibrium temperature occurs at 0°C and the steam-water equilibrium temperature occurs at 100°C.	Another scale frequently used in science is the Kelvin (Absolute) scale on which the ice-water equilibrium temperature occurs at 273 K and the steam-water equlibrium temperature occurs at 373 K.
IV. Phases of Matter	The term "phase" is used to refer to the gas, liquid, or solid form of matter.	The term "phase" is used instead of "state" to avoid confusion with other conditions such as "state of equilibrium."
	Change of phase of a substance is accompanied by the absorption or release of heat.	*Heating curves should be constructed in which the temperature is plotted against the time during which heat is added to a substance at a constant rate. Students should be able to interpret both heating and cooling curves.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
		 This curve, if read from right to left, would illustrate a cooling curve. *See also Unit 5*, V, p. 35.*
A. Gases	Gases take the shape and volume of the container.	
1. Boyle's law	At constant temperature, the volume of a given mass of gas varies inversely with the pressure exerted on it.	It is useful to introduce the equation $PV = k$ as an introduction to constants. *Students should be able to predict the direction of the change in volume with a specified change in pressure at constant temperature. Minimum requirements will be limited to simple mathematical relationships.*
2. Charles' law	At constant pressure, the volume of a given mass of gas varies directly with the Kelvin (Absolute) temperature.	The Kelvin (Absolute) temperature scale has its zero point at -273°C with the size of the degrees the same as on the Celsius scale. The volume of a gas decreases by 1/273 of its volume at 0°C for each decrease of 1°C, provided the pressure remains constant. While the basis for the Kelvin (Absolute) scale could be extrapolated from data relating to Charles' law experiments, it should be pointed out that absolute zero has not been reached, and that all gases liquefy before that point. *Students should be able to predict the direction of the change in volume with a specified change in temperature at constant pressure. Minimum requirements will be limited to simple mathematical relationships.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
3. Standard temperature and pressure (STP)	Standard temperature and pressure (STP) of a gas are defined as 0°C (273 K) and 760 mm of mercury (760 torr), or 1 atmosphere pressure.	Except for gases, standard temperature is usually 25°C (298 K). Since the volume of a given mass of gas varies with change in temperature and pressure, gas volumes are usually calculated to an arbitrary standard, STP.
4. Partial pressures	The pressure exerted by each of the gases in a gas mixture is called the partial pressure of that gas.	The total pressure of a gas mixture is equal to the sum of the individual partial pressures of the gases comprising the mixture.
5. Kinetic theory	Study of gas behavior has led to a model based on the following assumptions: • A gas is composed of individual particles which are in continuous, random, straight-line motion. • Collisions between gas particles may result in a transfer of energy between particles, but the net total energy of the system remains constant. • The volume of the gas particles themselves is ignored in comparison with the volume of the space in which they are contained. • Gas particles are considered as having no attraction for each other.	*In explaining and interpreting observed behavior, it is often convenient to use a model, which may be a "picture," a mathematical expression, or some other mechanism. Such models can be useful in the study of the behavior of gases. It should be emphasized that the model is only an approximation and is only as good as its ability to predict behavior under new conditions.* Not all of the particles of a gas have the same kinetic energy, but the average kinetic energy is proportional to the measured Kelvin (Absolute) temperature of the gas.
6. Deviations from the gas laws	Deviations from the gas laws occur because the model is not perfect; the gas particles do have volume and do exert some attraction for each other.	A gas which would conform strictly to the model would be an ideal gas. However, the model does not exactly represent any gas under all conditions. No real gas is ideal under all conditions of temperature and pressure.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	These factors become significant when the space between gas particles is reduced, as in conditions of relatively high pressure and low temperature.	*Experimental data (graphs) may be presented to show deviations from predicted values. Students should be encouraged to speculate as to the cause of these deviations. (Refer to Unit 3, van der Waals forces, p. 17.)*
7. Avogadro's hypothesis	Equal volumes of all gases under the same conditions of temperature and pressure contain equal numbers of particles.	For example, at the same temperature and pressure, the number of particles in 1 liter of hydrogen is the same as the number of particles in 1 liter of oxygen although the individual particles of oxygen are heavier and larger than the individual particles of hydrogen.
	The amount of matter that contains 6.02×10^{23} (Avogadro's number) particles is called a mole of matter. One mole of any substance contains as many particles (molecules, atoms, ions, etc.) as there are atoms of ^{12}C in 12.000g of ^{12}C isotope.	Since it is inconvenient to work with individual particles (atoms, molecules, ions, electrons, etc.), chemists have chosen a unit containing many particles for comparing amounts of different materials. The mole is a unit which contains 6.02×10^{23} particles.
	A mole of particles of any gas occupies a volume of 22.4 liters at STP.	
B. Liquids	Liquids have definite volume but take the shape of the container.	Particles of a liquid have no regular arrangement and are in constant motion.
1. Vapor pressure	When a liquid substance changes to a gas the process is called evaporation. Evaporation tends to take place at the surface of a liquid and at all temperatures.	The term "vapor" is frequently used to refer to the gas phase of a substance that is normally a liquid or solid at room temperature.
	In a closed system the vapor (gas) produced exerts a pressure which increases as the temperature of the liquid is raised and is specific for each substance and temperature.	*Demonstrations of the vapor pressure of a few liquids at various temperatures should be shown. The vapor pressure of water at various temperatures is given in the Reference Tables for Chemistry.*
2. Boiling point	A liquid will boil at the temperature at which the vapor pressure equals the pressure on the liquid.	
	The normal boiling point is the temperature at which the vapor pressure of the liquid equals one atmosphere.	Usually when reference is made to the "boiling point" of a substance, it is the normal boiling point that is indicated.

Topics	Understandings and Fundamental Concepts	Supplementary Information
3. Heat of vaporization	The energy required to vaporize a unit mass of a liquid at constant temperature is called its heat of vaporization.	The energy involved in the change of phase is required to overcome the forces of attraction between particles and does not increase their average kinetic energy. Thus there is no increase in temperature during this phase change. *For the purpose of the Regents examinations, heat of vaporization at the normal boiling temperature will be used. See also Unit 5*, IV, A, p. 35.*
C. Solids	Solids have definite shape and volume.	
1. Crystals	All true solids have crystalline structure. Crystals contain particles arranged in a regular geometric pattern. Particles are constantly vibrating even in the solid phase.	Certain materials often considered solids are really super-cooled liquids; for example, glass, some plastics. In solids, although the particles are vibrating, they do not change their relative positions in the regular geometric pattern.
2. Melting point	The normal melting point is the temperature at which a solid substance will change to a liquid at 1 atmosphere pressure.	Melting points can be determined from cooling curves which are obtained experimentally. A melting point may also be defined as the temperature at which the solid and liquid phases can exist in equilibrium. *See also Unit 5*, IV, B, p. 35.*
3. Heat of fusion	The energy required to change a unit mass of a solid to a liquid at constant temperature is called its heat of fusion.	
4. Sublimation	Sublimation is a change from the solid phase directly to the gas phase without passing through an apparent liquid phase.	*For the purpose of Regents examinations this definition of Sublimation will be used.* Solids that sublime have high vapor pressures and low intermolecular attractions. Examples of solids that sublime at room temperature are solid carbon dioxide (dry ice), solid iodine, and naphthalene.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Atoms.		
A. Introduction to atomic structure	Our concept of the nature of the atom has undergone change and will probably continue to do so.	*Teachers should acquaint pupils with at least a brief history of the development of the theory of atomic structure.*
B. Important subatomic particles	The atom is a complex unit of various particles.	
1. Electrons	An electron has a mass of approximately 1/1836 of a proton and a unit negative charge.	
2. Nucleons	The particles which compose the nucleus are called nucleons.	
a. Protons	A proton has a mass of approximately one atomic mass unit and a unit positive charge.	An atomic mass unit is defined as exactly 1/12 the mass of the ^{12}C atom.
b. Neutrons	A neutron has a mass of approximately one atomic mass unit and zero charge.	Although protons and neutrons are the only nuclear particles that have been identified in an intact nucleus, other particles have been identified among the break-down products of certain nuclear disintegrations. The relationship of these particles to the structure and stability of the nucleus is the subject of much current research.
C. Structure of atoms	Atoms differ in the number of protons and neutrons in the nucleus and in the configuration of electrons surrounding the nucleus.	
1. "Empty space" concept	Most of the atom consists of empty space.	Rutherford's gold foil experiments indicated the atom to be mostly empty space with the size of the nucleus very small compared to the size of the atom.
2. Nucleus	The mass of the atom is concentrated almost entirely in the nucleus.	The nature of the forces holding nuclear particles together is not adequately understood and is the subject of much current research.
a. Atomic number	The atomic number indicates the number of protons in the nucleus.	*Teachers may wish to discuss Moseley's work on the X-ray spectra in relation to the determination of the atomic numbers. Moseley's experiments measured the charge on the nucleus in units of elementary charge, which is interpreted as the number of protons in the nucleus.*

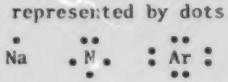
Topics	Understandings and Fundamental Concepts	Supplementary Information
b. Isotopes	Isotopes are atoms with the same atomic number but a different number of neutrons. For a given element the number of protons in the nucleus remains constant, but the number of neutrons may vary.	The atomic number identifies the element. The difference in the number of neutrons affects the mass of the atom.
c. Mass number	The mass number indicates the total number of protons and neutrons.	Since the masses of the protons and neutrons are each approximately one, the mass number approximates the mass of the isotope. The number of neutrons in an atom can be calculated by subtracting the atomic number from the mass number.
d. Atomic mass (weight)	The mass of a neutral atom, atomic mass, is measured in atomic mass units based on ^{12}C equal to 12.000 atomic mass units.	
	The atomic mass of an element given in the Reference Tables for Chemistry is the weighted average mass of the naturally occurring isotopes of that element. This average is weighted according to the proportions in which the isotopes occur.	Most elements occur naturally as mixtures of isotopes. This accounts for fractional atomic masses found in reference tables. In general, the mass number of the most abundant isotope of an element can be determined by rounding off the atomic mass of the element to the nearest whole number. The atomic mass (weight) of a single atom (isotope) such as neon-20 is 19.992 amu while the atomic mass (weight) of the element neon (which is the weighted average of all its natural isotopes) is 20.183 amu.
	The gram atomic mass (the mass of one mole of atoms) of an element is the mass in grams of Avogadro's number of atoms of that element as it occurs naturally.	The gram atomic mass is numerically equal to the atomic mass. See also Unit 5, Section II, A, p. 28.
3. Electrons	The electrons are outside the nucleus at various energy levels. In a neutral atom the total number of electrons is equal to the number of protons in the nucleus.	
D. Atomic structure models	The model for atomic structure of the elements has passed through many stages of development. In the Bohr model, electrons were considered to revolve around the nucleus in one of several concentric circular orbits.	*Models should be used to represent atomic structure. It should be pointed out to the student that these models are approximations and do not picture the actual atom. Students should be able to represent probable structures of atoms and ions and to indicate electronic charges using the model.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Principal energy levels	The principal energy levels can be denoted by the letters K, L, M, N, O, P, Q, or by the numbers 1, 2, 3, 4, 5, 6, 7.	The principal energy level approximates how far the electron is from the nucleus.
	Electrons in orbits near the nucleus are at lower energy levels than those in orbits farther from the nucleus.	When the electrons are in the lowest available energy levels the atom is said to be in the "ground state."
	When atoms absorb energy, electrons may shift to a higher energy level.	When electrons have absorbed energy and shifted to higher energy levels, the atom is said to be in an "excited state."
	The excited state is unstable, and the electrons fall back to lower energy levels.	
2. Quanta	Electrons can absorb energy only in discrete amounts called quanta.	
	When electrons return to a lower energy level, energy is emitted in quanta.	
3. Spectral lines	When electrons in an atom in the excited state return to lower energy levels, the energy is emitted as radiant energy of specific frequency, producing characteristic spectral lines which can be used to identify the element.	The study of spectral lines has provided much of the evidence regarding energy levels within the atom.
E. Orbital model of the atom	The orbital model differs from the Bohr model in that it does not represent electrons as moving in planetary orbits around the nucleus.	Although Bohr's model accounted for the lines of the hydrogen spectrum, it did not account for the spectra of heavier and more complicated atoms.
	The term orbital refers to the average region of most probable electron location. Electrons occupy orbitals that may differ in size, shape, or orientation in space.	
1. Energy levels	The energy levels of electrons within an atom are represented by quantum numbers.	
a. Principal quantum numbers	The principal quantum number (n) represents the principal energy level.	The principal quantum number (n) is equal to the number of the principal energy level as referred to under the Bohr atom and is the same as the period number in the periodic table.
b. Sublevels	The principal energy levels may be divided into sublevels.	Additional spectral lines appearing in the spectrum of atoms heavier than hydrogen can be explained only by assuming that the principal energy levels are divided into sublevels.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	The total number of possible sublevels for each principal energy level is equal to the number of the principal energy level. These sublevels are designated by the letters s, p, d, and f.	Principal energy level number 1 is made up of one energy sublevel called 1s; principal energy level 2 is made up of two sublevels called 2s and 2p; principal energy level 3 is made up of three sublevels called 3s, 3p, and 3d; principal energy level 4 is made up of four sublevels called 4s, 4p, 4d, and 4f. Within a given principal energy level the lowest in energy is the "s" sublevel and the highest is the "f" sublevel. Both the number of the principal energy level and one of the letters s, p, d, or f are used to describe the energy of an electron in a particular sublevel. (The number of occupied sublevels does not exceed four even when n is greater than 4.)
c. Orbitals	Each sublevel may consist of one or more orbitals with each orbital having a different spatial orientation. Each electron occupies an orbital. An orbital can hold no more than two electrons. The number of orbitals within the same principal energy level (n) is equal to n^2 when n is less than or equal to 4. The s sublevel consists of 1 orbital; the p sublevel consists of 3 orbitals; the d sublevel consists of 5 orbitals; the f sublevel consists of 7 orbitals.	The maximum number of electrons possible in the various energy levels and their distribution are shown in the following table.

Principal quantum number (n)	Number of orbitals (n^2)	s orbitals, p orbitals, d orbitals, f orbitals	Maximum number of electrons ($2n^2$)
1	1	1	2
2	4	1 3	8
3	9	1 3 5	18
4	16	1 3 5 7	32

2. Electron configurations	Electron configurations of the atoms in order of their atomic numbers starting with hydrogen can be built up by adding one electron at a time according to the following rules: • No more than two electrons can be accommodated in any orbital. • The added electron is placed in the unfilled orbital of lowest energy.	The two electrons in an orbital have opposite spins.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	• In a given sublevel, a second electron is *not* added to an orbital until each orbital in the sublevel contains one electron.	
	• No more than four orbitals are occupied in the outermost principal energy level of any atom.	In an electron configuration, the number of electrons in a sublevel is indicated by a superscript following the designation of the sublevel. For example, calcium would be represented as $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$ or [Ar] $4s^2$. Sulfur would be represented as $1s^2 2s^2 2p^6 3s^2 3p^4$ or [Ne] $3s^2 3p^4$. *For the purpose of Regents examinations, all electron configurations given in the Reference Tables for Chemistry may be used.*
3. Valence electrons	The electrons in the outermost principal energy level of an atom are referred to as the valence electrons.	The chemical properties of an atom are related to the valence electrons. The term "kernel" is sometimes used to refer to the atom exclusive of the valence electrons. The kernel consists of the nucleus and all electrons except the valence electrons. The valence electrons may be represented by electron dot symbols in which the kernel of the atom is represented by the letter symbol for the element and the valence electrons are represented by dots. For example: 
F. Ionization energy	Ionization energy is the amount of energy required to remove the most loosely bound electron from an atom in the gaseous phase.	Ionization energy refers to the removal of the first electron. The second ionization energy refers to the removal of the second most loosely bound electron. Each successive ionization energy is greater than the previous one. In the Reference Tables for Chemistry, ionization energies are expressed in kcal/mole of atoms. Ionization energies are used to compare chemical properties of elements. (See Unit 4, Section II, C and D, p. 22.)
II. Natural Radioactivity	Radioactivity is the spontaneous disintegration of the nucleus of an atom with the emission of particles and/or radiant energy.	When one element is changed to another element as a result of change in the nucleus, the change is called transmutation.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	Some naturally occurring elements are radioactive.	There are no known stable isotopes with atomic numbers greater than 83.
A. Difference in emanations	Emanations differ from each other in mass, charge, penetrating power, and ionizing power.	
	Nuclear disintegration of naturally radioactive atoms produces alpha particles, beta particles and gamma radiations.	*Students should be able to balance nuclear equations.*
1. Alpha decay	When an alpha particle is given off as the result of nuclear disintegration the reaction is called alpha decay.	An atom which emits an alpha particle is called an alpha emitter. When an atom emits an alpha particle, the atomic number is reduced by 2 and the mass number is reduced by 4. Example: $^{226}_{88}Ra \rightarrow {}^{222}_{86}Rn + {}^{4}_{2}He$
	Alpha particles are helium nuclei.	
2. Beta decay	When a beta particle is given off as the result of nuclear disintegration, the reaction is called beta decay.	An atom which emits a beta particle is called a beta emitter. When an atom emits a beta particle the atomic number is increased by 1 and the mass number remains the same. Example: $^{234}_{90}Th \rightarrow {}^{234}_{91}Pa + {}^{0}_{-1}e$
	Beta particles are high-speed electrons.	
3. Gamma radiation	Gamma rays are similar to high energy X-rays.	Gamma rays are not particles and do not have mass or charge. *The symbols used in nuclear chemistry are given in the Reference Tables for Chemistry.*
B. Separating emanations	Separation of emanations is possible by an electric field or magnetic field.	In an electric field, alpha particles are deflected toward the negative electrode, beta particles toward the positive electrode, while gamma rays are not affected by the field.

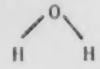
Topics	Understandings and Fundamental Concepts	Supplementary Information
C. Detection of radioactivity	The study of radioactivity is made possible by its ionizing, fluorescent, and photographic effects.	*Teachers may wish to discuss the use of photographic films, electroscopes, Geiger counters, or Scintillation counters, for the detection of radioactivity.*
D. Half-life	The half-life of a radioactive isotope is the time required for one-half of the nuclei of a given sample of that isotope to disintegrate.	*Simple examples should be used to demonstrate that the mass of an isotope remaining after a given time can be calculated.* *The half-lives of selected radioisotopes are given in the Reference Tables for Chemistry.*

(Questions based on this material may appear in Parts I and/or II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. The Nature of Chemical Bonding	A chemical bond results from the simultaneous attraction of electrons to two nuclei.	
A. Energy stored in bonds	Chemical energy is potential energy. The energy is stored in chemical bonds.	*Teachers may wish to use the analogy of potential energy stored in an object when it is lifted against the forces of gravity.*
B. Energy changes in bonding	When a chemical bond is formed, energy is released. When a chemical bond is broken, energy is absorbed.	When two atoms are held together by a chemical bond, they are generally at a lower energy condition than when they are separated.
C. Bonding and stability	Generally, systems at low energy levels are more stable than systems at high energy levels. Generally, chemical changes will occur among atoms if the changes lead to a lower energy condition, and hence a more stable structure.	For example, when an element such as fluorine, with a nearly filled outer shell, reacts with another element it usually releases energy in the process because the resulting state of the compound is at a lower energy level. Such a reaction would be an example of an exothermic reaction. When the compound thus formed is decomposed, energy must be put into the compound to decompose or break it up. Therefore, this would be an example of endothermic reaction. Exactly the same amount of energy would be required to force this endothermic reaction to proceed as was released when the compound was formed.
D. Electronegativity	Electronegativity is a measure of the ability of an atom to attract the electrons that form a bond between it and another atom.	Electronegativity values are based on an arbitrary scale on which fluorine, the most electronegative element, is assigned a value of 4.0. The electronegativity of an element does not measure the reactivity of the element.
II. Bonds between Atoms	The electrons involved in bond formation may be transferred from one atom to another, or may be shared equally or unequally between two atoms.	When atoms of the elements enter into chemical reaction, they do so in a manner that results in their becoming more like "inert" gas atoms in that they contain their maximum complement of valence electrons. This would place them in a condition of lowest energy content and maximum stability.

Topics	Understandings and Fundamental Concepts	Supplementary Information
A. Ionic	An ionic bond is formed by the transfer of one or more electrons from metals to nonmetals.	The transfer of electrons results in the formation of ions. In ionic bonding the number of electrons transferred is such that the atoms involved achieve an "inert" gas configuration. [Not generally true for transition elements.] Since the ion has a different electron configuration than the atom, the properties of the ion differ from those of the atom.
	Ionic bonds may also form between ions that were formed in a previous reaction.	Example: $AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
	Ionic solids have high melting points. As solids they do not conduct electricity.	In the geometric structure of the solid ionic crystal, ions form the crystal lattice and are held in relatively fixed positions by electrostatic attraction. When melted or dissolved in water, the crystal lattice is destroyed and the ions move freely. This free movement of ions permits electrical conductivity.
		Examples of ionic solids are sodium chloride and magnesium oxide.
B. Covalent	A covalent bond is formed when two atoms share electrons.	
1. Nonpolar covalent	When electrons are shared between atoms of the same element, they are shared equally and the resulting bond is nonpolar.	An example of a nonpolar covalent bond is found in the fluorine molecule:
2. Polar covalent	When electrons are shared between atoms of different elements, they are shared unequally (even though they may have the same apparent electronegativity values) and the resulting bond is polar.	An example of a polar covalent bond is found in the hydrogen chloride molecule:
3. Coordinate covalent	When the two shared electrons forming a covalent bond are both donated by one of the atoms, this bond is called a coordinate covalent bond.	A coordinate covalent bond, once formed, is not different from an ordinary covalent bond. The difference lies in the source of the electrons involved in the bond. The coordinate covalent bond is frequently involved in the bonding within polyatomic ions. The coordinate covalent bond is of importance in modern acid-base theories. (See Unit 7, Section III, B, p. 43).

Topics	Understandings and Fundamental Concepts	Supplementary Information
	The ionic or covalent *character* of a bond can be approximated from differences in electronegativity of the reacting species.	Electronegativity differences give an *indication* of the character of the bond. Electronegativity differences of 1.7 or greater indicate a bond that is predominantly ionic in character. Differences less than 1.7 indicate that the bond is predominantly covalent. (Some exceptions to this may be found. For example, the metal hydrides, with an electronegativity difference of less than 1.7, are predominantly ionic.)
4. Molecular substances	A molecule may be defined as a discrete particle formed by covalently bonded atoms.	A molecule has also been defined as the smallest particle of an element or compound capable of independent motion. Stable molecules usually have structures such that each atom has the electron configuration of an inert gas atom when the shared electrons forming each covalent bond are counted as belonging to both atoms connected by that covalent bond. Examples of molecules are H_2, NH_3, H_2O, HCl, CCl_4, S_8, $C_6H_{12}O_6$. *Teachers may wish to point out that there are exceptions to the octet rule in certain molecular species.*
	Molecular substances may exist as gases, liquids or solids, depending on the attraction that exists between the molecules. Generally, molecular solids are soft, electrical insulators, poor heat conductors, and have low melting points.	The forces of attraction between molecules and the general properties of the resulting substances are developed below under Section III of this unit, p. 17.
5. Network solids	Certain solids consist of covalently bonded atoms linked in a network which extends throughout the sample with an absence of simple discrete particles. Such a substance is said to be a network solid. Generally, network solids are hard, electrical insulators, poor conductors of heat, and have high melting points.	Examples of network solids are graphite (C), diamond (C), silicon carbide (SiC), and silicon dioxide (SiO_2).

Topics	Understandings and Fundamental Concepts	Supplementary Information
C. Metallic	Metallic bonding occurs between atoms of metals which have vacant valence orbitals and low ionization energies.	A metal consists of an arrangement of positive ions which are located at the crystal lattice sites and are immersed in a "sea" of mobile electrons. These mobile electrons can be considered as belonging to the whole crystal rather than to individual atoms. This mobility of electrons distinguishes the metallic bond from an ionic or covalent bond.
III. Molecular Attraction	Groups of atoms covalently bonded in a molecule may in turn be attracted to similar molecules or to ions.	
A. Dipoles	The asymmetric distribution of electrical charge in a molecule gives rise to a molecule which is polar in nature and is referred to as a dipole.	A molecule composed of only two atoms will be a dipole if the bond between the atoms is polar. Molecules composed of more than two atoms may be nonpolar, even though the individual bonds are polar, if the shape of the molecule is such that symmetric distribution of charge results. For example, CO_2 is a nonpolar molecule that can be represented as O = C = O while H_2O is a polar molecule that may be represented as  The symmetrical shape (tetrahedral) of the CH_4 molecule makes it nonpolar.
B. Hydrogen bonding	Hydrogen bonds are formed between molecules when hydrogen is covalently bonded to an element of small atomic radius and high electronegativity. This is a special case of dipole-dipole attraction.	When a hydrogen atom is bonded to a highly electronegative atom, the hydrogen has such a small share of the electron pair that it is almost like a bare proton. As such it can be attracted to the more electronegative atom of an adjacent molecule. This accounts for the high boiling point of H_2O as compared with the boiling point of H_2S. Hydrogen bonding is important in compounds of hydrogen with fluorine, oxygen, or nitrogen.
C. Van der Waals forces	Even in the absence of dipole attraction and hydrogen bonding, as in nonpolar molecules, weak attractive forces exist between molecules. These forces are called van der Waals forces.	Van der Waals forces make it possible for species of small nonpolar molecules (such as hydrogen, helium, oxygen, etc.) to exist in the liquid and solid phases under conditions of low temperature and high pressure.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	Van der Waals forces increase with increasing molecular size and with decreasing distance between the molecules.	The effect of molecular size on the magnitude of the van der Waals forces accounts for the increasing boiling points of a series of similar compounds (such as the alkane series of hydrocarbons). The increasing boiling points of the halogens with increasing molecular size can also be explained in terms of increasing van der Waals forces. Van der Waals forces appear to be due to chance distribution of electrons resulting in momentary dipole attractions.
D. Molecule-ion attraction	Polar solvents, when interacting with ionic compounds, attract ions from these compounds and form a solution.	Ionic compounds are generally soluble in polar solvents such as water, alcohol, and liquid ammonia. The negative ion of the substance being dissolved is attracted to the positive end of the adjacent polar molecules, while the positive ion is attracted to the negative end of the polar molecules. Water is the polar substance most commonly used to dissolve these ionic compounds. When an ionic compound is dissolved in water, its crystal lattice is destroyed and water molecules surround each ion forming hydrated ions. It is because water is a dipole that this attraction between the water molecules and the positive or negative ion exists. The orienting of water molecules around ions is called hydration of the ions. This process is important in aqueous chemistry.
IV. Directional Nature of Covalent Bonds	Generally the geometric structure of covalent substances which results from the directional nature of the covalent bond helps to explain properties of the resulting molecule.	The polarity of a water molecule is explained by the asymmetrical shape of the molecule.
V. Chemical Formula	A chemical formula is both a qualitative and a quantitative expression of the composition of an element or a compound.	
A. Symbol	A symbol may represent one atom or one mole of atoms of an element.	One mole of atoms contains Avogadro's number (6.02×10^{23}) of atoms.

Topics	Understandings and Fundamental Concepts	Supplementary Information
B. Formula	A formula is a statement in chemical symbols which represents the composition of a substance.	*Where necessary, teachers should review the use of subscripts, brackets, and coefficients.*
1. Molecular	A molecular formula indicates the total number of atoms of each element needed to form the molecule.	
2. Empirical	An empirical formula represents the simplest ratio in which the atoms combine to form a compound.	Empirical formulas are used to represent ionic compounds which do not exist as discrete molecular entities.
VI. Naming and Writing Formulas of Chemical Compounds	The chemical name of a compound generally indicates the chemical composition of the substance.	
	In binary compounds composed of metals and nonmetals, the metallic element is usually named and written first. The name of the nonmetal ends in *-ide*.	*Students should be able to determine the chemical formulas of common binary compounds from their names, and their names from the formulas.*
	In compounds composed of two nonmetals, the less electronegative element is usually named and written first. The name of the compound ends in *-ide*.	
	Prefixes are used to indicate the number of atoms of each nonmetal in the compound.	Examples include carbon dioxide, sulfur trioxide, carbon tetrachloride, etc.
		Students should also be familiar with the Stock system of naming compounds of two nonmetals. For example, N_2O is named nitrogen (I) oxide, NO is named nitrogen (II) oxide, and NO_2 is named nitrogen (IV) oxide.
	In the naming of other types of compounds the metal is usually named first and the polyatomic negative ion is named last.	*Students should be able to use the selected polyatomic ions, given in the Reference Tables for Chemistry, in the writing of formulas and naming of compounds.*
		In naming compounds of metals which may have more than one oxidation number, the Stock system should be used. In this system, Roman numerals indicate the oxidation number of the metal ion. For example, FeO is named iron (II) oxide, and Fe_2O_3 is named iron (III) oxide.

Topics	Understandings and Fundamental Concepts	Supplementary Information
		Students should be familiar with the naming of acids, bases, and salts containing polyatomic ions given in the Reference Tables for Chemistry.
VII. Chemical Equations	An equation represents the qualitative and quantitative changes in bonding and energy that take place in a chemical reaction.	*Simple equations that can be balanced by inspection should be introduced; for example:* $2H_2 + O_2 \rightarrow 2H_2O + heat$
		Balancing redox equations is taken up in Unit 8, Section III, p. 49 and Unit 8, Section VI, p. 67.*
	Equations must conform to the laws of conservation of mass and charge.	Energy is usually omitted from the equation when we are not concerned with it.
		In an equation, it is often desirable to indicate the phase of reactants and products. This may be done by using the symbols (s) - solid; (ℓ) - liquid; (g) - gas; and (aq) - in aqueous solution. The equation shown above could be written: $2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell) + heat$

This material need not be taught as a unit, but may be incorporated in various places in the syllabus at the discretion of the teacher.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Development of Periodic Table	The periodic table of the elements has passed through many stages of development, evolving into the present form.	Observed regularities in properties of elements led Mendeleev and others to consider these regularities to be functions of the atomic mass. Moseley established that properties of elements are periodic functions of the atomic number.
	The atomic number is the basis of the arrangement in the present form of the periodic table.	
	The properties of the elements depend on the structure of the atom, and vary with the atomic number in a systematic way.	
II. Properties of Elements in the Periodic Table	The horizontal rows of the periodic table are called periods, or rows. The properties of elements change systematically through a period.	The number of a period is the same as the number of the outermost principal energy level containing electrons.
	The vertical columns of the periodic table are called groups or families. The elements of a group exhibit similar or related properties.	
A. Covalent atomic radius	The covalent atomic radius is one-half the measured internuclear distance in the solid phase. The covalent atomic radius is a periodic property of the elements.	The relation between covalent atomic radius and atomic number can be interpreted in terms of the arrangement of electrons in the oribitals of atoms and in terms of nuclear charge.
	Within a single period of the periodic table, the covalent atomic radius generally decreases as the atomic number increases.	Within any one period the electrons in the outer orbitals are arranged around a kernel containing the same number of filled levels. As one proceeds from left to right in the period, increase in nuclear charge due to the increasing number of protons pulls the electrons more tightly around the nucleus. This increased attraction more than balances the repulsion between the added electron and other electrons, and thus the covalent atomic radius is reduced.
	The members of any group in the periodic table generally show an increase in covalent atomic radius with an increase in atomic number.	For a group of elements, the atoms of each successive member have a larger kernel containing more filled levels. Hence the electrons in the unfilled orbitals are farther from the nucleus. This results in an increase in covalent atomic radius as the atomic number increases among the elements in a group.

Topics	Understandings and Fundamental Concepts	Supplementary Information
		Measurement of inter-nuclear distances is difficult and different methods may give slightly different results. For purposes of examination, the data found in the Reference Tables for Chemistry should be used.
B. Ionic radius	A loss or gain of electrons by an atom causes a corresponding change in size.	Metal atoms lose one or more electrons when they form ions. Ionic radii of metals are smaller than the corresponding atomic radii. Nonmetal atoms gain one or more electrons when they form ions. Ionic radii of nonmetals are larger than the corresponding atomic radii. Atomic and ionic radii are usually measured in Angstrom (Å) units (1Å = 10^{-10} meter).
C. Metals	Metal atoms possess relatively low ionization energies. Metal atoms tend to lose electrons to form positive ions when combining with other elements. Metals usually possess the properties of high thermal and electrical conductivities, metallic luster, malleability, and ductility.	More than two-thirds of the elements are metals. Metallic properties are most pronounced in those elements on the lower left side of the periodic table. All metals except mercury are solids at room temperature.
D. Nonmetals	Nonmetal atoms possess high ionization energies and have high electronegativities. Nonmetal atoms tend to gain electrons when in combination with metals, or to share electrons when in combination with other elements. Nonmetals in the solid phase tend to be brittle, to have low thermal and electrical conductivities, and to lack metallic luster.	Nonmetallic properties are most pronounced in those elements in the upper right corner of the periodic table. Nonmetals tend to be gases, molecular solids, or network solids. (The exception is bromine which is a volatile liquid at room temperature.)
E. Metalloids	Metalloids are those elements which have some properties characteristic of metals and other properties characteristic of nonmetals.	Examples of metalloids are boron, silicon, arsenic, and tellurium.

Topics	Understandings and Fundamental Concepts	Supplementary Information
III. Chemistry of a Group (Family)	The elements in the periodic table are divided into Groups A and B except for Group 0.	*Minimum requirements for the detailed study of individual groups are indicated in the following subsections of this section. It is expected that teachers will discuss the properties of the elements and their important compounds in sufficient detail to illustrate the principles outlined in these subsections. The properties of the compounds studied should be related to uses in home and industry.*
	The elements in each group exhibit related chemical properties. Similarities in chemical properties within a group are associated with similarity in the number of valence electrons.	Related chemical properties are illustrated by the similarity in the type of compound formed by the members of a group. For example, the elements in Group IA form chlorides having the general formula MCl where M represents any member of the group. Elements in Group IIA form chlorides having the general formula MCl_2.
	In general, the properties of elements in a group change progessively as the atomic number increases.	Properties of the members of a group should be explained in terms of bonding, electronegativity, atomic size, and so forth. It should be pointed out that anomalies in the properties of elements within a group do occur. For example, in Group IIIA, boron does not form an ion as do other members of the group. These anomalies occur most frequently among the elements in Period 2 because of the relative closeness of the valence electrons to the nucleus and the relatively small shielding effect of the two electrons in the 1s sublevel.
	In the "A" groups of the periodic table, as the atomic number increases: • The radius of the atom increases. • The ionization energy of the element generally decreases.	The ionization energy decreases with increasing atomic number because of the increased distance of the valence electron(s) from the nucleus and also because of the increased shielding effect produced as new occupied energy levels are added to the atom.

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	• The electronegativity of the element generally decreases.	It should be pointed out that the relative tendency of the atoms to form compounds (sometimes called the reactivity of the element) cannot be predicted from the electronegativity of the atom, but should be ascertained from the Table of Standard Electrode Potentials given in the Reference Tables for Chemistry.
	• The elements tend to have more metallic properties.	
A. Groups IA and IIA	Groups IA and IIA include the most reactive metals.	Elements in Group IA are called the alkali metals. *Note: Hydrogen is <u>not</u> an alkali metal.* Elements in Group IIA are called the alkaline earth metals. Because of their reactivity, Groups IA and IIA elements occur in nature only in compounds.
	The elements in both groups have relatively low ionization energies and electronegativities. They lose electrons readily to form ionic compounds that are relatively stable.	The elements in these groups cannot achieve an inert gas structure by covalent bonding, and form only ionic compounds.
	Generally, the reactivity within both groups tends to increase with an increase in atomic number.	Exceptions to this occur in Group IA. The anomaly of lithium is due to the small size of the lithium atom and ion.
	In the same period, each Group IA metal is more reactive than the corresponding Group IIA metal.	
	The elements in both groups are usually reduced to their free state by the electrolysis of their fused compounds.	
B. Groups VA and VIA	The elements in Group VA show a marked progression from nonmetallic to metallic properties with increasing atomic number.	Nitrogen and phosphorus are typical nonmetals; arsenic is classified as a metalloid; bismuth is metallic in both appearance and properties.
	The element nitrogen is relatively inactive at room temperature.	In general, the reactivity of nonmetals in the same group decreases with increasing atomic number.

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Topics	Understandings and Fundamental Concepts	Supplementary Information
		Nitrogen exists as a diatomic molecule with a triple bond between the two atoms. The high energy required to break the triple N≡N bond explains the relative inactivity of nitrogen.
	Nitrogen compounds are essential constituents of all living matter.	The instability of many nitrogen compounds makes them useful as explosives.
	Generally, nitrogen compounds are relatively unstable.	
	The element phosphorus is more reactive than nitrogen at room temperature.	Phosphorus does not exist as a diatomic molecule at room temperature, but exists as a tetratomic molecule, P_4.
	Phosphorus compounds are essential constituents of all living matter.	*Students who have studied biology will be familiar with calcium phosphate as a constituent of bones and teeth and with the phosphate linkage present in DNA and RNA.*
	The elements in Group VIA show a marked progression from nonmetallic to metallic properties with an increase in atomic number.	Oxygen and sulfur are typical nonmetals; tellurium is classified as a metalloid; polonium shows metallic properties.
	The element oxygen is an active nonmetal.	Oxygen forms compounds with most elements. The existence of oxygen in its free state, in spite of its high reactivity, is explained by the continuous production of oxygen by plants during photosynthesis. Because of its high electronegativity, oxygen in compounds always shows a negative oxidation state except when combined with fluorine.
	Sulfur is less reactive than oxygen.	Sulfur in compounds shows both negative and positive oxidation states.
		Examples of elements having allotropic forms are phosphorus, oxygen and sulfur.
	Selenium and tellurium are rare elements.	Selenium and tellurium show a negative oxidation state when combined with hydrogen. In most other compounds they show positive oxidation states.
	Polonium is a radioactive element.	Polonium is an alpha emitter. Polonium is a degradation product of uranium.

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C. Group VIIA	The elements in Group VIIA are typical nonmetals.	Group VIIA is known as the halogen family. Although the metallic character increases with increasing atomic number, none of the elements in the group is a metal. (Astatine is radioactive with a short half-life. It has not been found in nature and its properties are not well known.)
	The elements in Group VIIA have relatively high electronegativities.	Fluorine has the highest electronegativity of any element and in compounds can show only a negative oxidation state. The other elements of the group may exhibit positive oxidation states in combination with more electronegative elements. The ease with which positive oxidation states of the halogens are formed increases with increasing atomic number.
	The physical form of the free element at room temperature varies with increasing atomic number.	At room temperature fluorine and chlorine are gases, bromine is a liquid and iodine a solid. The change in physical form is due to the increase in van der Waals forces. (See Unit 3, Section III, C., p. 17.)
	The elements are usually prepared from the corresponding halide ion by removing one of the electrons from the ion.	Because of their high reactivity, the halogens occur in nature only in compounds. Since fluorine is the most electronegative element, there is no chemical oxidizing agent that can oxidize the fluoride ion to fluorine. Fluorine is prepared by the electrolysis of its fused compounds. Chlorine, bromine and iodine can be prepared by various chemical methods.
	The ions of transition elements are usually colored, both in solid compounds and in solution.	
D. Group 0	Group 0 elements are monatomic gases. The atoms of these elements have complete outer shells, which result in an electron configuration that is stable.	Group 0 elements are referred to by a variety of terms including rare gases, "inert" gases, and noble gases. The term "inert" is no longer strictly applicable to this group, since it is possible to form compounds of krypton, xenon, and radon with fluorine and oxygen.

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		However, the term is still in general use and the electron configuration is quite generally referred to as the "inert gas structure".
E. Transition elements	Transition elements are those elements in which electrons from the two outermost sublevels may be involved in a chemical reaction. Because of this, these elements generally exhibit multiple positive oxidation states.	The transition elements are found in the "B" groups and Group VIII of the periodic table.
IV. Chemistry of a Period	A study of the Group A elements in a period, from left to right leads to certain generalizations. In each period, as the atomic number increases: • The radius of the atom generally decreases. • The ionization energy of the element generally increases. • The electronegativity of the element generally increases. • The elements generally change from very active metals to less active nonmetals to very active nonmetals and, finally, to an "inert" gas. • There is a transition from positive to negative oxidation states. • The metallic characteristics of the "A" group elements decrease.	Elements near the center of the period may exhibit both positive and negative oxidation states.

This material need not be taught as a unit, but may be incorporated in various places in the syllabus at the discretion of the teacher.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Mole Interpretation	A mole contains Avogadro's number of particles.	See also Unit 1, Section IV, A, 7, p. 5.
II. Use of the Mole Concept	The mole may be used in calculations involving the number of particles (atoms, molecules, ions, electrons, or other particles) involved in chemical reactions, the mass of elements or compounds, or the volume relationships in gases.	
A. Gram atomic mass (gram-atom)	Gram atomic mass (gram-atom) of an element represents the mass in grams of Avogadro's number of atoms of the element.	See also Unit 2, Section I, C, 2d, p. 8.
		The gram atomic mass is numerically equal to the atomic mass as shown in the periodic table.
B. Gram molecular mass	The gram molecular mass (mass of 1 mole) is the sum of the gram atomic masses of the atoms that make up a particular molecule.	
	The gram formula mass is the sum of the gram atomic masses of the atoms that make up a particular formula.	The gram formula mass calculated from the empirical formula is used for ionic substances and network solids, since they are not molecular substances.
C. Molar volume of a gas	A mole (Avogadro's number) of molecules of any gas occupies a volume of 22.4 liters at STP. It has a mass equal to the molecular mass expressed in grams.	See also Unit 1, Section IV, A, 7, p. 5.
III. Stoichiometry	Stoichiometry is the study of the quantitative relationships implied by chemical formulas and by chemical equations.	In stoichiometry it is frequently convenient to use the mole interpretation and mole relationships in the solving of problems.

Topics	Understandings and Fundamental Concepts	Supplementary Information
A. Problems involving formulas		
1. Percent composition	The percent by mass of an element in a compound can be calculated from the ratio of the mass of that element represented in the formula to the formula mass of the compound.	Sample problems: 1. Calculate the percent composition by mass of water (H_2O). 2. Calculate the percent, by mass, of water of hydration in gypsum ($CaSO_4 \cdot 2H_2O$).
2. Empirical formula	An empirical formula represents the simplest ratio in which the atoms combine to form a compound.	*Students should be able to get simple empirical formulas from molecular formulas and from gram-atom ratios.* The empirical formula of H_2O_2 is HO. The empirical formula of a compound that has a ratio of 1.0 gram-atom of carbon to 2.0 gram-atoms of hydrogen is CH_2. *Students should also be able to determine simple molecular formulas from empirical formulas.* If the empirical formula of a compound is CH, the molecular formula could be any simple multiple of CH such as C_2H_2 and C_3H_3.
B. Problems involving equations	The coefficients used in balancing a chemical equation represent moles. In stoichiometric problems involving equations it is assumed that the reaction is a single reaction (with no side reactions), that the reaction goes to completion, and that the reactants are completely consumed.	*In solving many problems involving equations it is convenient to use mole relationships. However, other methods may be used.*
1. Mass problems	A balanced equation shows the mole proportions of products and reactants. It is possible to determine the mass of one substance that reacts with, or is produced from, a given mass of another substance.	Equation: $2NaCl \rightarrow 2Na + Cl_2$ The equation indicates that 2 moles of NaCl will produce 2 moles of Na and 1 mole of Cl_2. Sample problem: How many grams of sodium chloride are needed to produce 355 grams of chlorine gas?

Topics	Understandings and Fundamental Concepts	Supplementary Information
2. Mass - volume problems	Since, in the balanced equation, the mole unit serves to relate the quantities of products and reactants, it is possible to determine quantitative results in desired units, which may not necessarily be the same as the original units.	Sample problem: Given the reaction: $Zn + 2HCl \rightarrow ZnCl_2 + H_2$, how many grams of zinc metal are needed to react with hydrochloric acid to produce 11.2 liters of hydrogen gas at STP?
3. Volume problems	Since one mole of any gas occupies the same volume as one mole of any other gas, at the same temperature and pressure, the volumes of gases involved in a reaction are proportional to the number of moles indicated by numerical coefficients in a balanced equation.	When reacting gases are at the same temperature and pressure, the volume ratio is equal to the mole ratio (Avogadro's hypothesis). Sample problem: In the reaction: $N_2 + 3H_2 \rightarrow 2NH_3$, calculate the volume of hydrogen required to form 100 liters of NH_3.
IV. Solutions	A solution is a homogeneous mixture of two or more substances, the composition of which may vary within limits. The component of a solution which is usually a liquid and is present in excess is called the solvent, while the other component is called the solute.	Most solutions dealt with in beginning courses in chemistry are aqueous solutions.
A. Methods of indicating concentrations	Concentration of solutions may be indicated in a variety of ways.	*Students should be familiar with the terms miscible and immiscible; dilute and concentrated; saturated, unsaturated, and supersaturated; and with the use and interpretation of the solubility charts in the Reference Tables for Chemistry.*
1. Molarity	The molarity (M) of a solution is the number of moles of solute contained in a liter (1000 ml) of solution. $M = \frac{\text{number of moles of solute}}{\text{1 liter of solution}}$	A two molar (2M) solution contains 2 moles of solute per liter of solution and 0.1 molar solution (0.1M) contains 0.1 mole of solute per liter of solution. The concentration in moles per liter multiplied by the volume in liters equals the number of moles of solute in the solution. Moles of solute = molarity x volume in liters. The mass in grams of solute can be determined by multiplying the number of moles of solute by the mass of 1 mole.

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Grams of solute = no. of moles x mass of 1 mole.

Sample problems:

1. How many moles of NaOH are contained in 200 ml of a 0.1M solution of NaOH?

2. How many grams of NaOH are contained in 500 ml of a 0.5M solution of NaOH?

3. What is the molarity of a solution of KOH if 500 ml of the solution contains 5.6 grams of KOH?

(Questions based on this material may appear only in Part II of the Regents examination.)

Teachers who elect this "Unit 5*" may wish to teach materials outlined here in appropriate places in the syllabus.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. The Mole - Additional Problems		Students should be able to calculate: • The number of molecules contained in a specific volume of a gas at STP. Sample problem: What is the number of molecules contained in 5.6 liters of nitrogen gas at STP? • The number of molecules in a given mass of a substance. Sample problem: What is the number of molecules contained in 14 grams of nitrogen gas? • The number of atoms of a particular element in a given number of moles of the substance. Sample problem: How many atoms of hydrogen are contained in 0.50 mole of NH_3 gas? Students should also be able to use equations and calculate: • The number of grams of a substance that can be reacted with or produced from a known number of moles of a reactant. Sample problem: Given the reaction $N_2 + O_2 \rightarrow 2\ NO$, what is the total number of grams of oxygen needed to react completely with 2.0 moles of nitrogen?

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
		• The volume of a gas at STP that can be reacted with or produced from a known number of moles of a reactant. Sample problem: Given the reaction $N_2 + 3H_2 \rightarrow 2NH_3$, how many liters of NH_3 gas at STP can be produced from 1.0 mole of N_2? • The number of molecules that can be reacted with or produced from a known number of moles of a reactant. Sample problem: Given the reaction $2 H_2 + O_2 \rightarrow 2H_2O$, how many molecules of hydrogen are needed to produce 6.0 moles of water? Some additional problems: Given the reaction: $Zn + 2HCl \rightarrow ZnCl_2 + H_2$, how many molecules of H_2 are produced by the reaction of 6.54 grams of Zn with excess dilute HCl? Given the reaction: $2C_2H_6 + 7 O_2 \rightarrow 4 CO_2 + 6 H_2O$, how many liters of CO_2 at STP are produced by the combustion of 22.4 liters of C_2H_6 at STP?
II. Formula from Percent Composition	The empirical formula of a compound can be determined from the percent composition of the compound and the atomic masses of the elements. If the compound is molecular, the molecular formula can also be determined if the gram molecular mass is known.	Sample problems: A compound was found by analysis to consist of 80% carbon and 20% hydrogen by mass. What is the empirical formula of the compound? The gram molecular mass of the compound was found to be 30 grams. What is the molecular formula of the compound?

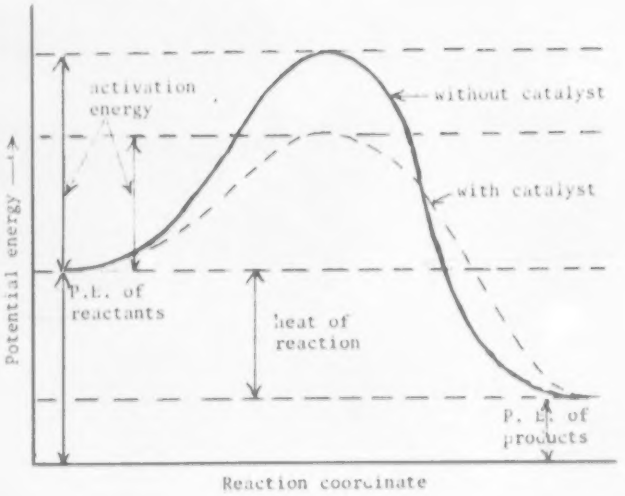
(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
III. Gram Molecular Mass from Gas Density	The molecular mass of a gas can be determined from the gas density (mass per unit volume, usually expressed in grams/liter) by using the relationship between the molar volume and the mass of one mole of the gas.	Sample problem: The density of a gas is 1.35 grams/liter at STP. Calculate the gram molecular mass of the gas.
IV. Effect of Solute on Solvent	The presence of dissolved particles affects some properties of the solvent.	
	Properties which depend on the relative number of particles rather than on the nature of the particles are called colligative properties.	Colligative properties as related to solutions include changes in boiling point, freezing point, vapor pressure, and osmotic pressure. *Changes in vapor pressure and osmotic pressure due to the presence of solute particles in the solvent will not be included on Regents examinations.*
A. Boiling point elevation	The presence of a nonvolatile solute raises the boiling point of the solvent by an amount that is proportional to the concentration of dissolved solute particles.	One mole of particles per 1000 grams of water raises the boiling point of water 0.52°C. The relationship between moles of solute and grams of solvent is expressed in the concentration unit, molality. The molality (m) of a solution is the number of moles of a solute dissolved in 1000 grams of solvent.
B. Freezing point depression	The presence of a solute lowers the freezing point of the solvent by an amount that is proportional to the concentration of dissolved solute particles.	One mole of particles per 1000 grams of water lowers the freezing point of water 1.86°C. *The molal boiling point elevation constant and the molal freezing point depression constant for water are given in the Reference Tables for Chemistry.* *Teachers may wish to point out that measurement of the boiling point elevation or freezing point depression of a solvent can be used to determine the molecular mass of the solute.*

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
C. Abnormal behavior of electrolytes	An ionic solute increases the boiling point and lowers the freezing point of a solvent more than a nonelectrolyte of equal molality.	This behavior of electrolytes in solution gives credence to the existence of ions.
V. Calorimetry	The energy changes involved in chemical reactions are measured in calorimeters.	*For the purpose of examination, the number of calories of energy gained or lost in a reaction equals the mass of water in grams times the change in temperature in °C. The heat gained or lost by the calorimeter itself is ignored. The heat capacity of water (in the liquid phase) is assumed to to be 1 calorie per gram per Celsius degree.*
A. Heat of vaporization - problems	For definition, see Unit 1, IV, B, 3, p. 6.	*Teachers may wish to have students determine heats of vaporization experimentally. Students should be able to use the formula:* *Number of calories of energy absorbed during vaporization at the boiling point equals the mass of water in grams times the heat of vaporization.*
B. Heat of fusion - problems	For definition, see Unit 1, IV, C, 3, p. 6.	*Teachers may wish to have students determine heats of fusion experimentally. Students should be able to use the formula:* *Number of calories of energy absorbed during fusion at the melting point equals the mass of ice in grams times the heat of fusion.*
VI. Combined Gas Laws	Since changes in volume, pressure, and temperature often occur simultaneously, it is convenient to combine the two equations of Boyle and Charles into a single equation: $\frac{PV}{T}$ = constant	The combined gas law equation may be written: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, where P_1, V_1, and T_1 are the original conditions of pressure, volume and Kelvin temperature, and P_2, V_2, and T_2 are the corresponding values of the final conditions.
VII. Graham's Law	Under the same conditions of temperature and pressure, gases diffuse at a rate inversely proportional to the square roots of their molecular masses.	*Calculations using the diffusion formula for Graham's law will not be included on the Regent's examination.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Kinetics	Chemical kinetics is the branch of chemistry concerned with the rate of chemical reactions and the mechanisms by which chemical reactions occur.	The rate of a chemical reaction is measured in terms of the number of moles of reactant consumed (or moles of product formed) per unit volume in a unit of time.
	The mechanism of a chemical reaction is the sequence of stepwise reactions by which the overall change occurs.	Even though many reactions take place as a result of a series of steps, only the net reaction is often observable. The net reaction represents a summation of all the changes that occur.
A. Role of energy in reactions	Energy is required to initiate a chemical reaction. Energy may be released or absorbed in a chemical reaction.	By graphing the potential energy of the reactants, the activation energy, and the potential energy of the products against a time sequence, one may describe the energies involved in a chemical reaction.
1. Activation energy	Activation energy is the minimum energy required to initiate a reaction.	
2. Heat (enthalpy) of reaction	Heat (enthalpy) of reaction (ΔH) is the heat energy released or absorbed in the formation of the products. It represents the difference in heat content between the products and the reactants. $\Delta H = H_{products} - H_{reactants}$	
	In an exothermic reaction, energy is released, the products have a lower potential energy than the reactants, and the sign of ΔH is negative. In an endothermic reaction, energy is absorbed, the products have a higher potential energy than the reactants, and the sign of ΔH is positive.	The sign that may be used when energy is included in a chemical equation should not be confused with the sign for ΔH. For example, the equation for the reaction of hydrogen and oxygen to form water may be written: $H_2(g) + \frac{1}{2}O_2(g) = H_2O(\ell) + 68.3$ kcal Since this is an exothermic reaction, the sign of ΔH is negative and its value is -68.3 kcal. In equations that include heat, the phase of each species should be specified.

Topics	Understandings and Fundamental Concepts	Supplementary Information
3. Potential energy diagram	For a given reaction, the activation energy and heat of reaction can be shown graphically in a potential energy diagram by plotting potential energy against a reaction coordinate representing the process of the reaction. The difference in potential energy between the final products and the initial reactants represents the heat (enthalpy) of reaction. If the potential energy of the products is lower than the potential energy of the reactants, energy has been liberated (exothermic reaction). If the potential energy of the products is higher than the potential energy of the reactants, energy has been absorbed (endothermic reaction). The highest point in the curve represents the potential energy of the intermediate products (activated complexes). The difference between this point and the initial potential energy of the reactants represents the activation energy of the reaction.	*Students should be able to identify the heat of reaction and activation energy, and to distinguish between exothermic and endothermic reactions from a potential energy diagram.*  **The graph shown above, when read from left to right, represents an exothermic reaction with the products at a lower potential energy than the reactants. If the graph is reversed, or read from right to left, it would represent an endothermic reaction.**
8. Factors affecting rate of reaction	Chemical reactions depend on collisions between the reacting particles, atoms, molecules, or ions. The rate of reaction is affected by the number of collisions occurring and the fraction of these collisions that are effective.	*In considering the factors affecting the rate of a reaction it is useful to use a model of colliding particles.* *Teachers may elect to discuss examples of the applications of principles affecting rates of reactions to commercial processes.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Nature of the reactants	Since bonds may be broken or formed in a reaction, the nature of the bond is an important factor affecting reaction rates.	Reactions that involve negligible bond rearrangements are usually rapid at room temperature; for example, reactions of ionic substances in aqueous solutions. Reactions that involve the breaking of bonds tend to be slow at room temperature; for example, the reaction between hydrogen and oxygen.
2. Concentration	An increase in the concentration of one or more reactants generally increases the rate of reaction.	An increase in the concentration of a reactant increases the frequency of collisions. Concentrations are usually measured in moles per liter.
	In a gaseous system, an increase in pressure will result in an increase in concentration, and thus an increase in the rate of reaction.	The effect of pressure on the rate of reaction should not be confused with the effect of pressure on chemical equilibrium. (See Section II, C, 1, b of this unit, p. 41.)
3. Temperature	An increase in temperature increases the rate of all chemical reactions.	An increase in temperature increases the speed (and thus the kinetic energy) of the particles, and increases not only the number of collisions per unit time, but of greater importance, the effectiveness of the collisions.
4. Surface area	Increasing the surface area of reactants increases their rate of reaction.	In heterogeneous reactions, surface area plays an important role. For example, a given amount of zinc will react more readily with dilute hydrochloric acid if the surface area of zinc is increased by using smaller pieces.
5. Catalysts	Catalysts change the activation energy required and thus change the rate of reaction.	A catalyst changes the mechanism of a reaction to one involving less activation energy, but does not change the overall process.
	A catalyst does not initiate a chemical reaction.	*For examination purposes, interpret the term "catalyst" to be a substance whose presence increases the rate of a chemical reaction.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
II. Equilibrium	Most reactions are reversible.	
	Equilibrium is a state of balance between two opposing reactions (physical or chemical) occurring at the same rate.	Equilibrium is dynamic and only describes the overall appearance of the system. It does not describe the activity of individual particles. The word dynamic implies motion, and dynamic equilibrium is that condition in which the interaction of the particles of the reactants in one direction is balanced by the interaction of the particles of the products in the opposite direction. Although the reaction rates for the opposing reactions are equal, a state of equilibrium may exist in which the quantities of reactants and products are not equal. Thus equilibrium may be reached when only a small quantity of the products has been formed or when only a small quantity of reactants remains. For a system in equilibrium, a change in conditions (such as temperature, concentration, or pressure) may result in a change in the equilibrium point. Because the reactions in an equilibrium are reversible, it follows that equilibrium may be attained either from the forward or the reverse reaction.
A. Phase equilibrium	In general, phase changes (solid to liquid, or liquid to gas) are reversible and, in a closed system, equilibrium may be attained.	In general, if a solid or a liquid is confined in a closed container, eventually there will be enough particles in the vapor phase so that the rate of return is equal to the rate of escape. Thus, a dynamic equilibrium results in which there is an equilibrium vapor pressure characteristic of the solid or the liquid.
B. Solution equilibrium		
1. Gases in liquids	In a closed system, equilibrium may exist between a gas dissolved in a liquid and the undissolved gas above the liquid.	

Topics	Understandings and Fundamental Concepts	Supplementary Information
	The equilibrium between dissolved and undissolved gas is affected by temperature and pressure.	Increased temperature decreases the solubility of gases in liquids. Increased pressure increases the solubility of gases in liquids.
2. Solids in liquids	A solution equilibrium exists when the opposing processes of dissolving and crystallizing of a solute occur at equal rates. A solution exhibiting equilibrium between the dissolved and undissolved solute must be a saturated solution.	
3. Solubility	The solubility of a solute is defined as the maximum mass of that solute dissolved in a given volume of solvent under specified conditions. Solubility may be defined as the concentration of solute in a saturated solution.	*Students should be able to interpret the solubility curves given in the Reference Tables for Chemistry. See also Unit 5, IV, A, p. 30 and Unit 6, p. 45.*
C. Chemical equilibrium	Chemical equilibrium is attained when the concentration of the reactants and products remains constant.	When observable changes (such as color, pressure, and temperature) no longer occur in a reacting chemical system, the system has reached a state of equilibrium. At this point the forward reaction and the reverse reaction are occurring at equal rates.
1. Le Chatelier's principle	If a stress, such as a change in concentration, pressure, or temperature, is applied to a system at equilibrium, the equilibrium is shifted in a way that tends to relieve the effects of the stress.	When a chemical system at equilibrium is disturbed, chemical reaction occurs and equilibrium is reestablished at a different point (i.e., with new concentrations of reactants and products). *Minimum requirements will be limited to a qualitative treatment of the application of Le Chatelier's principle to systems in equilibrium.*
a. Effect of concentration	Increasing the concentration of one substance in a reaction at equilibrium will cause the reaction to go in such a direction as to consume the increase. Eventually a new equilibrium will be established.	For example, in the equilibrium reaction, $N_2 + 3H_2 = 2NH_3$, increasing the concentration of either nitrogen or hydrogen will increase the rate of ammonia formation. If the system remains closed, the increased concentration of ammonia that results will increase the rate of decomposition of ammonia, and a new equilibrium point will be established.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	Removal of one of the products of a reaction results in a decrease in its concentration, and will cause the reaction to go in such a direction as to increase the concentration of the product.	Removal of a product tends to cause the forward reaction to go more nearly to completion. Continuous removal of the product may destroy the equilibrium system by removing all of that substance necessary for the reverse reaction. Products may be removed from a reaction, wholly or in part, by the formation of a gas, the formation of an insoluble product (precipitate) or, in an ionic reaction, by the formation of an essentially un-ionized product such as water. Such reactions are said to go to "completion."
b. Effect of pressure	A change in pressure affects chemical equilibria in which gases are involved. An increase in pressure will displace the point of equilibrium in the direction that favors the formation of a lesser number of moles. If no change in the number of moles is involved in the reaction, a change of pressure has no effect on the equilibrium.	For example, in the Haber process, $N_2(g) + 3H_2(g) = 2NH_3(g)$, four moles of reactants form two moles of product. The forward reaction thus results in a decrease in the total number of moles. Thus an increase in pressure will result in an increased production of ammonia which relieves the effect of increased pressure. However, in the reaction $H_2(g) + Cl_2(g) = 2HCl(g)$, a change in pressure does not affect the equilibrium. *See also Unit 10*, p. 72.*
c. Effect of temperature	When the temperature of a system in equilibrium is raised, the equilibrium is displaced in such a way that heat is absorbed.	Chemical changes involve either the evolution or the absorption of energy. In every system in equilibrium, an endothermic and exothermic reaction are taking place simultaneously. The endothermic reaction is favored by an increase in temperature, the exothermic reaction by a decrease in temperature. It should be pointed out that the rates of all reactions, both endothermic and exothermic, are increased by an increase in temperature. (See Section I, B, 3 of this unit, p. 38.) However, the opposing reactions are increased unequally, resulting in a displacement of the equilibrium. For example, in the Haber process, $N_2(g) + 3H_2(g) = 2NH_3(g) + 22$ kcal, raising the temperature favors the decomposition of ammonia.

Topics	Understandings and Fundamental Concepts	Supplementary Information
d. Effect of catalyst	In a system in equilibrium, a catalyst increases the rate of both the forward and reverse reactions equally, and produces no net change in the equilibrium concentrations.	A catalyst may cause equilibrium to be reached more quickly, but does not affect the point of equilibrium.
2. Law of chemical equilibrium	When a reversible reaction has attained equilibrium at a given temperature, the product of the molar concentrations of the substances to the right of the equation, divided by the product of the molar concentrations of the substances to the left (each concentration raised to the power equal to the number of moles of that substance appearing in the equation), is a constant.	*Minimum requirements will be limited to predicting qualitatively the effects of concentration changes in simple reactions, and to the interpretation of the significance of the value of K.*
	For the reaction: $a\,M(g) + b\,N(g) = c\,P(g) + d\,Q(g)$, $\frac{[P]^c[Q]^d}{[M]^a[N]^b} = K$ = a constant at constant temperature	*The derivation of the law of chemical equilibrium will not be subject to examination.*
	This constant is called the equilibrium constant.	In the mathematical expression of this law, square brackets are used to indicate "concentrations measured in moles per liter." By convention, in calculating the equilibrium constant, the concentrations of products on the right of the chemical equation form the numerator and the concentrations of reactants on the left form the denominator.
	Usually the equilibrium constants for various reactions are expressed as K_a for acids, K_b for bases, K_{eq} for chemical reactions, K_{sp} for solubility of solids, and K_w for the ionization of water.	The concentration of a solid or a liquid is essentially constant. In the expression of the equilibrium constant for any reaction involving a solid or a liquid, the concentration of the solid or liquid can be included in the constant and, therefore, does not appear in the equation.
	The magnitude of K is used by chemists to predict the extent of chemical reactions.	The equilibrium constant has a numerical value for any given chemical reaction at a particular temperature. This value remains constant even though the concentrations of the substances involved may increase or decrease. In general, in the reaction given, an increase in the concentration of P would cause the reaction to go to the left, thus decreasing the concentration of Q and increasing the concentrations of M and N. The value of K would remain constant.
	A large value of K indicates the favoring of products; that is, the equilibrium mixture consists largely of products.	
	A small value of K indicates that reactants are favored.	

Topics	Understandings and Fundamental Concepts	Supplementary Information
	The equilibrium constant changes with a change of temperature.	In a reversible reaction the reaction rates of the forward and reverse reactions are not affected equally by change in temperature.
III. Spontaneous Reactions	Spontaneous reactions depend on the balance between two fundamental tendencies in nature: (1) toward a lower energy state; and (2) toward randomness.	
A. Energy changes	At constant temperature and pressure, a system tends to undergo a reaction so that, in its final state, it has lower energy than in its initial state.	A system tends to change from a state of high energy to one of low energy. For chemical systems this energy is represented mathematically as ΔH. (See Section I, A, 2 of this unit, p. 36.) This tendency in nature favors the exothermic reaction in which ΔH is negative.
B. Entropy changes	Entropy is a measure of the disorder, randomness, or lack of organization of a system.	The solid phase, in regular crystalline arrangement, is more organized than the liquid phase; the liquid phase is more organized than the gaseous phase.
	Entropy is so defined that the more random a system is, the higher the entropy.	An increase in entropy during a change in the state of a system means that in its final state the system is more disordered (random) than in its initial state.
	High entropy (randomness) is favored by high temperatures.	High temperatures increase the rate of motion of the particles, and thus tend to increase randomness.
	At constant temperature, a system tends to undergo a reaction so that in its final state it has higher entropy (greater randomness) than in its initial state.	A system tends to change from state of great order to a state of less order. For chemical systems this change in entropy is represented mathematically as ΔS.

(Questions based on this material may be tested only in Part II of the Regents examination.)

Teachers who elect this unit may wish to teach materials outlined here in the appropriate places in Unit 6.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Free Energy Change	The difference between energy change and entropy change is the free energy change (ΔG). $\Delta G = \Delta H - T\Delta S$	*Minimum requirements for Regent's examination will be limited to the interpretation of the significance of the various terms in the equation.*
II. Predicting Spontaneous Reactions	For a spontaneous change to occur in a system, the free energy change (ΔG) must be negative. In a system at equilibrium, the free energy change is zero.	
	When the two factors, tendency toward lower energy content and tendency toward higher entropy in a system, cannot be satisfied simultaneously, the spontaneous change that may take place will be determined by the factor that is dominant at the temperature of the system.	From energy changes alone, exothermic reactions would always be expected to occur spontaneously, and endothermic reactions would never be expected to occur spontaneously. Exceptions to both of these predictions may occur when a change in entropy opposes an exothermic reaction or favors an endothermic reaction. For example: (1) The change in phase from water to ice is an exothermic reaction and, from the consideration of energy only, water might be expected to freeze spontaneously at any temperature. However, the tendency toward higher entropy favors the reaction from ice to water. At temperatures below the freezing point the energy change is dominant and water will freeze spontaneously. At temperatures above the freezing point the entropy change becomes the dominant factor, and ice melts. (2) The reaction, $2KClO_3 \rightarrow 2KCl + 3O_2$ is an endothermic reaction, and the energy change would oppose a spontaneous reaction. However, the reaction results in an increase in entropy, due to the formation of a gas and a solid. At high temperatures the effect of the entropy changes becomes sufficient to overcome the effect of the energy change, and the reaction takes place. If the temperature is not high enough, the reaction will not take place spontaneously.

(Questions based on this material may be tested only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
III. Solubility Product Constant (K_{sp})	The solubility product constant, K_{sp}, is a measure of the concentration of slightly soluble salts in water. In a saturated solution of an ionic solid, an equilibrium is established between the ions in the saturated solution and the excess solid phase. For the reaction: $AB\ (s) = A^+(aq) + B\ (aq)$ $K_{sp} = [A^+]\ [B^-]$ The value of the K_{sp} changes with a change of temperature.	*Minimum requirements for examination will be limited to the qualitative interpretation of the significance of relative K_{sp} values.* The concentration (mass/unit volume) of a solid is essentially constant. In the expression of the equilibrium constant for any reaction involving a solid, the concentration of the solid can therefore be included in the constant and does not appear in the equation. Thus, in the case of the solubility equilibrium, application of the law of chemical equilibrium would give $K = \frac{[A^+]\ [B^-]}{[AB]}$ Since [AB] itself is constant, we may say $[A^+]\ [B^-] = K\ [AB] = K_{sp}$. The magnitude of K_{sp} is used in comparing the solubilities of slightly soluble salts. For example, at room temperature $K_{sp}\ CaSO_4 = 2.4 \times 10^{-5}$; $K_{sp}\ BaSO_4 = 1.6 \times 10^{-9}$. Since the K_{sp} of $BaSO_4$ is less than that of $CaSO_4$, $BaSO_4$ must be less soluble than $CaSO_4$, and would be precipitated at a lower concentration.
A. Common ion effect	The addition of a common ion to the solution of a slightly soluble salt results in a decrease in the solubility of the salt.	Consider the equilibrium: $AgCl_{(s)} = Ag^+_{(aq)} + Cl^-_{(aq)}$ Addition of NaCl to this equilibrium system increases the $Cl^-_{(aq)}$ concentration and shifts the equilibrium to the left. This results in a decrease in the solubility of AgCl.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Electrolytes	An electrolyte will dissolve in water to form a solution that will conduct an electric current.	All ionic compounds are electrolytes. Some polar covalent compounds form ions and conduct electricity when dissolved in water. Examples: HCl, HBr. See also p. 50 of this unit.
	The ability of a solution to conduct an electric current is due to the presence of ions that are free to move.	Arrhenius provided the first explanation of the behavior of electrolytes in aqueous solution.
	Weak electrolytes in aqueous solution attain an equilibrium between ions and the undissociated compound. The equilibrium constant for such systems is called the dissociation constant.	See also Unit 5*, IV, p. 34.
	The value of the dissociation constant changes with a change of temperature.	
II. Acids and Bases	Acids and bases may be defined in terms of operational definitions.	An operational definition is one based on experimental observations and includes a set of conditions to apply when a term is used in a particular situation.
	Conceptual definitions of acids and bases have been extended as the understanding of acid-base reactions has grown, and as principles of these reactions have been applied to reactions not in aqueous solutions.	A conceptual definition is based on the interpretation of observed facts.
A. Acids	Acids may be defined in terms of their characteristic properties.	These properties can be observed experimentally, and form the basis of the operational definition of an acid.
	• Aqueous solutions of acids conduct electricity.	Acids conduct electricity in relation to the degree of their ionization. A few acids ionize almost completely in aqueous solution and are strong electrolytes (strong acids), while others ionize only to a slight degree and are weak electrolytes (weak acids). (See the Reference Tables for Chemistry.)
	• Acids will react with certain metals to liberate hydrogen gas.	Acids will react with those metals below hydrogen in the Table of Standard Electrode Potentials given in Reference Tables for Chemistry.

Topics	Understandings and Fundamental Concepts	Supplementary Information
		Some acids, in addition to their acid properties, have strong oxidizing ability and thus, except in very dilute solution, do not release hydrogen gas on reaction with metals. For example, nitric acid and concentrated sulfuric acid have strong oxidizing properties.
	• Acids cause color changes in acid-base indicators.	Acid-base indicators are substances which have different colors in acid and basic solutions. Two common indicators are litmus which is blue in basic solution and red in acid solution, and phenolphthalein which is pink in basic solution and colorless in acid solution. Note that these indicators do not change color exactly at pH 7. Different indicators change color at different concentrations of hydrogen ions.
	• Acids react with hydroxides to form water and a salt.	When hydrogen ions react with hydroxide ions, water is formed. This reaction is called neutralization.
	• Dilute aqueous solutions of acids have a sour taste.	*Teachers should not permit students to taste laboratory chemicals, but should point out that the sour taste of many common foods is due to the presence of acids.*
1. Arrhenius theory	An acid is a substance that yields hydrogen ions as the only positive ions in aqueous solution. The characteristic properties of acids in aqueous solution are due to an excess of hydrogen ions.	This conceptual definition is adequate when considering reactions in aqueous solutions. As knowledge of the mechanism of chemical reactions has increased, more inclusive definitions have been advanced.
2. Brönsted-Lowry theory	An acid is any species (molecule or ion) that can donate a proton to another species.	The Brönsted-Lowry theory does not replace the Arrhenius theory, but rather extends it. The Brönsted-Lowry definition of an acid includes all substances that are acids according to the Arrhenius definition.

Topics	Understandings and Fundamental Concepts	Supplementary Information
		In addition, some molecules and ions are classified as acids under the Brönsted-Lowry definition that are not acids in the Arrhenius sense. For example, in the reaction, $NH_3 + H_2O = NH_4^+ + OH^-$, the water molecule donates a proton to the ammonia, and is considered an acid in the Brönsted-Lowry sense.
B. Bases	Bases may be defined in terms of their characteristic properties.	These properties form the basis of the operational definition of a base.
	• Aqueous solutions of bases conduct electricity.	Some examples of the relative degree of ionization are shown in the Reference Tables for Chemistry.
	• Bases cause color changes in acid-base indicators.	
	• Bases react with acids to form water and a salt.	
	• Aqueous solutions of bases feel slippery.	Strong bases have a caustic action on the skin.
1. Arrhenius theory	A base is a substance that yields hydroxide ions as the only negative ions in aqueous solution.	According to the Arrhenius definition, the only bases are hydroxides.
	The characteristic properties of bases in aqueous solution are due to the hydroxide ion.	
2. Brönsted-Lowry theory	A base is any species (molecule or ion) that can combine with a proton.	The Brönsted-Lowry definition extends the Arrhenius definition to include many species in addition to the OH^- that can accept a proton. For example, in the reaction, $H_2O + HCl = H_3O^+ + Cl^-$, the water molecule combines with a proton to form the hydronium ion, and is here considered a base in the Brönsted-Lowry sense.
C. Amphoteric (amphiprotic) substances	An amphoteric (amphiprotic) substance is one that can act either as an acid or as a base, depending on its chemical environment.	In the Brönsted-Lowry sense, species (for example, H_2O, HSO_4^-) which can either donate or accept a proton are amphiprotic.
III. Acid-Base Reactions		
A. Neutralization	Acid-base neutralization pertains to the reaction which occurs when equivalent quantities of an acid and a hydroxide are mixed.	One mole of hydrogen ions will react with one mole of hydroxide ions to form water. $H^+ + OH^- \rightarrow H_2O$ *Note: The spectator ions are not included in the equation.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Acid-base titration	The molarity of an acid or base of unknown concentration can be determined by slowly combining it with a base (or acid) of known molarity (standard solution) until neutralization occurs. This process is called titration. The end point of the titration may be determined by the use of appropriate indicators or by temperature changes or electrode potential changes.	The molarity of a solution of unknown concentration can be calculated from an understanding of the molar relationship involved. Although normality is no longer a suggested way to express concentration, it will probably be in common usage for some time. A limitation of the normality scale is that a given solution may be more than one normality, depending on the reaction for which it is used. *Minimum requirements for examination will be limited to simple calculations involving acid-base neutralization reactions.*
2. Salts	A salt is an ionic compound containing positive ions other than hydrogen and negative ions other than hydroxide. Some salts in aqueous solution react with the water to form solutions that are acidic or basic. This process is called hydrolysis.	Most salts are strong electrolytes and are considered to be completely dissociated in aqueous solution. Salts can be considered as having been derived from the neutralization of an acid and a base. Salts formed from strong acids and weak bases will give aqueous solutions that are acidic. Salts formed from weak acids and strong bases will give aqueous solutions that are basic. Salts formed from strong acids and strong bases do not hydrolyze in solution.
B. Conjugate acid-base pair	According to the Brönsted-Lowry theory, acid-base reactions involve a transfer of protons from the acid to the base. In an acid-base reaction, an acid transfers a proton to become a conjugate base. This acid and its newly formed base form a conjugate acid-base pair. A base gains a proton to become a conjugate acid, forming a second acid-base pair. Each pair, made up of an acid and its base, is related by the transfer of a proton. **The strongest acids have the weakest conjugate bases and the strongest bases have the weakest conjugate acids.**	In order to accept a proton, a base must have at least one pair of unshared electrons. The proton (H^+) will share a pair of electrons belonging to the base, forming a coordinate covalent bond. (See Unit 3, Section II, B, p. 15.) In the following reactions the two conjugate pairs ($acid_1$-$base_1$ and $acid_2$-$base_2$) are identified by subscripts. $Base_1 + Acid_1 = Acid_2 + Base_2$ $H_2O + HCl = H_3O^+ + Cl^-$ $NH_3 + H_2O = NH_4^+ + OH^-$ **Refer to the Reference Tables for Chemistry.**

Topics

IV. Ionization Constant

Understandings and Fundamental Concepts

The equilibrium constant for the ionization of acids (K_a) is a convenient method for comparing the relative strength of acids. For the reaction

$HB\ (aq) = H^+(aq) + B^-(aq)$,

$$K_a = \frac{[H^+]\ [B^-]}{[HB]}$$

A. K_w

In water and aqueous solutions the product of the hydrogen ion concentration and the hydroxide ion concentration is a constant at constant temperature. This constant, K_w, is useful in problems involving hydrogen ion and hydroxide ion concentrations.

$K_w = [H^+]\ [OH^-]$

Supplementary Information

Minimum requirements for examination will be limited to the interpretation of the significance of the value of K_a.

Ionization constants can be calculated for all acids that are not completely ionized. For acids that are completely ionized there is no equilibrium. The denominator [HB] approaches zero and K_a approaches infinity.

Ionization constants such as those given in the Reference Tables for Chemistry can be used in comparing the relative strength of acids.

For example, an acid with a $K_a = 1.8 \times 10^{-5}$ (although a weak acid) is stronger than an acid with $K_a = 1.0 \times 10^{-7}$.

Ionization constants such as those given in the Reference Tables for Chemistry can be used in comparing the relative strengths of acids.

Students should be able to solve problems based on the relationship

$K_w = [H^+]\ [OH^-] = 1.0 \times 10^{-14}$.

Since the concentration of water is essentially constant for all reactions in aqueous solution, its value can be included in the ionization constant giving

$K_w = [H^+]\ [OH^-] = 1.0 \times 10^{-14}$.

The numerical value of K_w at 25° C is 1.0×10^{-14}.

In pure water, $[H^+] = [OH^-] = 1.0 \times 10^{-7}$ at 25° C.

Topics	Understandings and Fundamental Concepts	Supplementary Information
B. pH	The pH of a solution indicates the concentration of hydrogen ions (acid strength) in a solution. A pH of 7 is neutral. A pH of less than 7 is acidic. A pH greater than 7 is basic.	pH is the logarithm of the reciprocal (negative logarithm) of the hydrogen ion concentration. $pH = \log \frac{1}{[H^+]} = -\log [H^+]$ $[H^+] = 10^{-pH}$ Pure water has a pH of 7 at 25°C. *Minimum requirements for examination will be limited to the calculation of molar hydrogen ion concentrations from pH values and calculation of pH values from molar hydrogen ion concentrations. Only calculations involving whole number pH values will be used.*

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Redox (Oxidation-Reduction)	Redox reactions result from the competition for electrons between atoms.	Redox is a term used for oxidation-reduction.
A. Oxidation	Oxidation represents a loss, or an apparent loss, of electrons. Oxidation refers to any chemical change in which there is an increase in oxidation number. The particle that increases in oxidation number is said to be oxidized. The particle that is being oxidized acts as a reducing agent.	
B. Reduction	Reduction represents a gain, or an apparent gain, of electrons. Reduction refers to any chemical change in which there is a decrease in oxidation number. The particle that decreases in oxidation number is said to be reduced. The particle that is being reduced acts as an oxidizing agent.	
C. Oxidation number	The oxidation number (oxidation state) of an atom is the charge which an atom has, or appears to have, when electrons are counted according to certain arbitrary rules. In assigning oxidation numbers, electrons shared between two unlike atoms are counted as belonging to the more electronegative atom; electrons shared between two like atoms are divided equally between the sharing atoms.	The oxidation number, although arbitrary, is a convenient notation for keeping track of the number of electrons involved in a chemical reaction. Application of the general rules results in the following operational rules for determining oxidation number: • In the free elements, each atom has an oxidation number of zero. For example, hydrogen in H_2, sodium in Na, and sulfur in S_8, all have oxidation numbers of zero. • In simple ions (ions containing one atom) the oxidation number is equal to the charge on the ion.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	All oxidation numbers must be consistent with the conservation of charge.	• In $CaCl_2$, calcium has an oxidation number +2 and chlorine -1; iron in $FeCl_2$ has an oxidation number of +2 while in $FeCl_3$ it has an oxidation number of +3. • All metals in Group IA form only 1+ ions and their oxidation number is +1 in all compounds. • All metals in Group IIA form only 2+ ions and their oxidation number is +2 in all compounds. • Oxygen has an oxidation number of -2 in all its compounds EXCEPT in peroxides (such as H_2O_2) when it is -1, and in compounds with fluorine when it may be +1 or +2. For example, in H_2SO_4, oxygen has an oxidation number of -2. • Hydrogen has an oxidation number of +1 in all its compounds (such as HCl and H_2SO_4) EXCEPT in the metal hydrides (such as LiH and CaH_2) when it is -1. *By convention, the charge of an ion like that of barium is written as 2+ whereas the oxidation number of barium is written as +2.* For neutral molecules the oxidation number of all the atoms must add up to zero. For example, in H_2SO_4 the two hydrogens contribute a total of +2, the four oxygens contribute a total of -8. Therefore, the sulfur must contribute an oxidation number of +6. For polyatomic ions (charged particles that contain more than one atom) the oxidation numbers of all the atoms must add up to the charge on the ion. For example, in SO_4^{2-}, the four oxygens contribute a total oxidation number of -8. Therefore, the sulfur must contribute an oxidation number of +6 to give the ion a charge of 2-.

Topics	Understandings and Fundamental Concepts	Supplementary Information
D. Redox reactions	Oxidation and reduction occur simultaneously; one cannot occur without the other.	In oxidation and reduction the increase and decrease of oxidation number results from a shift of electrons. The only way in which electrons can be shifted away from an atom (oxidation) is for them to be pulled toward another atom or ion (reduction)
	There is a conservation of charge as well as a conservation of mass in a redox reaction.	
II. Electrochemistry		
A. Half-reactions	A redox reaction may be considered in two parts, one representing a loss of electrons (oxidation), and the other representing a gain of electrons (reduction). Each reaction is known as a half-reaction.	A separate equation showing gain or loss of electrons (electronic equation) can be written for each half-reaction. For example: $Mg + Cl_2 \rightarrow MgCl_2$ can be represented as $Mg \rightarrow Mg^{2+} + 2e^-$ (oxidation) $Cl_2 + 2e^- \rightarrow 2Cl^-$ (reduction)
B. Half-cells	It is possible to set up reactions so that each half of a redox reaction takes place in a separate vessel, provided that the vessels are connected by an external conductor and a salt bridge or porous partition which permits the migration of ions but does not allow the solutions to mix.	
C. Chemical cells (Electrochemical cells)	Redox reactions that occur spontaneously may be employed to provide a source of electrical energy. When the two half-cells of a redox reaction are connected by an external conductor and a salt bridge, a flow of electrons (electric current) is produced.	In a chemical cell, a chemical reaction is used to produce an electric current.
D. Electrolytic cells	Redox reactions that do not occur spontaneously can be forced to take place by supplying energy with an externally applied electric current. The use of an electric current to bring about a chemical reaction is called electrolysis.	In an electrolytic cell, an electric current is used to produce a chemical reaction.

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Topics	Understandings and Fundamental Concepts	Supplementary Information
III. Balancing Simple Redox Equations	In any reaction the loss of electrons by the species oxidized must be equal to the gain of electrons by the species reduced. There is a conservation of charge as well as a conservation of mass in a redox reaction.	$Cu + 2Ag^{+} \rightarrow Cu^{2+} + 2Ag$ $2Al + 6H^{+} \rightarrow 2Al^{3+} + 3H_2$

(Questions based on this material may be tested only in Part II of the Regents examination.)

Teachers who elect this unit may wish to teach materials outlined here in the appropriate places in Unit 8.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Standard Electrode Potentials		
A. Half-cell potential	Comparison of the driving force of a half-reaction with that of the hydrogen standard establishes a scale of voltages. When each half-reaction is compared to the standard under specified conditions of concentration, temperature, and pressure, standard electrode potentials can be obtained. The standard electrode potential (E°) gives the potential difference in volts between the specified half-reaction and the hydrogen half-reaction.	It is impossible to measure the absolute reducing tendency of a half-reaction except by comparison with an attendant half-reaction. For purposes of measuring the relative reducing tendency of a half-reaction, it has been found convenient to adopt the half reaction for the reduction of the hydrogen ion, $2\ H^+ + 2e^- \rightarrow H_2$, as an arbitrary standard. A table of Standard Electrode Potentials is included in the Reference Tables for Chemistry.
B. Use of standard electrode potentials	Standard electrode potentials are useful in determining whether or not a specific redox reaction will take place. Any pair of half-reactions can be combined to give the complete reaction for a cell whose potential difference can be calculated by adding the appropriate half-cell potentials.	In combining half-reactions it must be remembered that, in any redox reaction, there must be an oxidation half-reaction and a reduction half-reaction. In the table of Standard Electrode Potentials, half-reactions are written as reductions. To obtain the equation and the potential for the oxidation half-reaction, the equation as written in the table must be reversed, and the sign of E° changed. For example, in the reaction: $Mg + Cl_2 \rightarrow MgCl_2$, the magnesium is oxidized. Therefore, to obtain the equation and the potential for this oxidation half-reaction, the equation for magnesium as written in the table must be reversed and the sign of E° changed. Chlorine is reduced in this reaction and, therefore, its reduction potential is obtained directly from the table.

(Questions based on this material may be tested only in Part II of the Regents examination.)

Topics

Understandings and Fundamental Concepts

In combining half-reactions, if the potential (E°) for the overall reaction is positive, the reaction is spontaneous.

Supplementary Information

The two half-reactions may now be combined:

$Mg \rightarrow Mg^{2+} + 2e^-$ $\quad E^\circ = +2.37$

$Cl_2 + 2e^- \rightarrow 2Cl^-$ $\quad E^\circ = +1.36$

$Mg + Cl_2 \rightarrow Mg^{2+} + 2Cl^-$ $\quad E^\circ = +3.73$

In combining half-reactions, the electron transfer must be balanced. For example, in the reaction

$2Na + Cl_2 \rightarrow 2NaCl$,

the two half-reactions would be combined as follows:

$2\,(Na \rightarrow Na^+ + e^-)$ $\quad E^\circ = +2.71$

$Cl_2 + 2e^- \rightarrow 2Cl^-$ $\quad E^\circ = +1.36$

$2Na + Cl_2 \rightarrow 2Na^+ + 2Cl^-$ $\quad E^\circ = +4.07$

NOTE: The Standard Electrode Potentials (E°) are not multiplied by the coefficients in calculating the E° for the reaction.

Metals with negative reduction potentials will produce hydrogen upon reaction with an acid.

For example, if magnesium metal is added to a solution of an acid, the net reaction, if one occurs, could be represented by the equation:

$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$.

Combining the half-reactions:

$Mg \rightarrow Mg^{2+} + 2e^-$ $\quad E^\circ = +2.37$

$2H^+ + 2e^- \rightarrow H_2$ $\quad E^\circ = 0.00$

$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$ $\quad E^\circ = +2.37$

Since this value is positive, we know the reaction will take place spontaneously.

(Questions based on this material may be tested only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
	In combining half-reactions, if the potential (E°) for the overall reaction is negative, a reaction will not take place spontaneously.	Metals with positive reduction potentials will not produce hydrogen on reaction with an acid. If copper metal is added to a solution of an acid, the reaction would not occur spontaneously. For example: $Cu + 2H^+ \rightarrow Cu^{2+} + H_2$. Combining the half-reactions: $Cu \rightarrow Cu^{2+} + 2e^-$ $\quad E^\circ = -0.34$ $2H^+ + 2e^- \rightarrow H_2$ $\quad E^\circ = 0.00$ $Cu + 2H^+ \rightarrow Cu^{2+} + H_2$ $\quad E^\circ = -0.34$ Since the resulting E° is negative, we know the reaction will not take place spontaneously.
C. Equilibrium	Equilibrium is attained in chemical cells when the voltage measured is equal to zero.	The E° values given in the table of Standard Electrode Potentials are for definite concentrations. As a reaction proceeds, these concentrations change and the measured value falls off until at equilibrium the measured voltage is equal to zero.
II. Chemical Cells - Calculations		*Students should be able to:* • *Predict the direction of electron flow* • *Predict the direction of ion movement* • *Calculate the net potential for the redox reaction.* Commercial batteries are examples of chemical cells.

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(Questions based on this material may be tested only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
III. Electrolytic Cell Reactions		Electrolysis of water, electrolysis of concentrated brine, and electrolysis of molten salts are examples of the use of an electric current to produce a chemical reaction. $2\ H_2O$ + electricity → $2H_2$ + O_2 $2\ NaCl + 2\ H_2O$ + electricity → $2\ NaOH + H_2 + Cl_2$ $2\ KCl$ (fused) + electricity → $2\ K + Cl_2$
IV. Electrodes		*For the purpose of the Regents examination, the definitions as stated in this section must be used. The definitions for the terms "Cathode" and "Anode" are based on information found in the "Handbook of Chemistry and Physics" published by CRC Press, Cleveland, Ohio.*
A. Cathode	The electrode at which reduction occurs in a cell is called a cathode. In a chemical cell, the cathode is the positive electrode. In an electrolytic cell, the cathode is the negative electrode.	
B. Anode	The electrode at which oxidation occurs in a cell is called an anode. In a chemical cell, the anode is the negative electrode. In an electrolytic cell, the anode is the positive electrode.	
V. Electroplating	In this process an electric current is used to produce a chemical reaction. The material to be plated is the cathode and the metal used for plating is the anode. The electrolyte used is a suitable salt of the metal used for plating.	

(Questions based on this material may be tested only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
VI. Balancing Redox Equations	In any reaction the loss of electrons by the species oxidized must be equal to the gain of electrons by the species reduced. There is a conservation of charge as well as a conservation of mass in a redox reaction.	*Several methods have been used to balance redox reactions. Teachers may use the method they prefer.* One method is illustrated below for balancing the reaction between aqueous nitric acid and solid iodine. Given the unbalanced equation: $HNO_3 + I_2 \rightarrow HIO_3 + NO_2 + H_2O$ Step 1. Assign oxidation numbers to each element. $\overset{+1\ +5\ -2}{H\ N\ O_3} + \overset{0}{I_2} \rightarrow \overset{+1\ +5\ -2}{H\ I\ O_3} + \overset{+4\ -2}{N\ O_2} + \overset{+1\ -2}{H_2\ O}$ Step 2. Determine the change in oxidation number (transfer of electrons) of the elements: for nitrogen, +5 → +4 (gain of 1 electron); for iodine, 0 → +5 (loss of 5 electrons). Step 3: Write partial electronic equations for the materials oxidized and reduced. $N^{+5} + 1e^- \rightarrow N^{+4}$ $I_2^0 \rightarrow 2I^{+5} + 10e^-$ Step 4. Balance the electrons gained and lost by writing appropriate coefficients for the two half-reactions. $10\ (N^{+5} + 1e^- \rightarrow N^{+4})$ $1\ (I_2^0 \rightarrow 2I^{+5} + 10e^-)$ $10\ N^{+5} + I_2 \rightarrow 10\ N^{+4} + 2\ I^{+5}$ This equation is the net equation and does not include the spectator ions or atoms that are not involved in the redox reaction. $10\ HNO_3 + I_2 \rightarrow 2\ HIO_3 + 10\ NO_2 + H_2O$ Step 5. Insert other coefficients consistent with the conservation of matter. Balance by inspection. $10\ HNO_3 + I_2 \rightarrow 2\ HIO_3 + 10\ NO_2 + 4\ H_2O$

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Definitions	Organic chemistry is the chemistry of the compounds of carbon.	Organic compounds occur extensively in nature. All living things are composed predominantly of organic compounds.
	Carbon is able to form four covalent bonds not only with other kinds of atoms but also indefinitely with other carbon atoms. This makes possible a very large number of compounds.	Organic compounds are more numerous than inorganic compounds.
		The major sources of raw materials from which organic chemicals are obtained are petroleum, coal, wood and other plant products, and animal sources.
II. Characteristics of Organic Compounds	Organic compounds are generally insoluble in water and soluble in nonaqueous solvents.	Organic compounds are generally nonpolar and tend to dissolve in nonpolar solvents. Those organic compounds that are somewhat polar, such as acetic acid, are soluble in water.
	Organic compounds are generally nonelectrolytes.	Organic acids are weak electrolytes.
	Organic compounds generally have low melting points.	Since most organic compounds are essentially nonpolar, the intermolecular forces are weak. Thus, the compounds have relatively low melting points (under 300°C).
	Reactions involving organic compounds are generally slower than those involving inorganic compounds.	Because of strong covalent bonding within the molecule, organic compounds do not readily form activated complexes (intermediates), and thus, reactions take place slowly. The activation energy required for organic reactions is generally high.
A. Bonding	The carbon atom usually forms compounds by covalent bonding.	
	The carbon atom has four valence electrons and can form four covalent bonds.	The four single bonds of the carbon atom are spatially directed toward the corners of a regular tetrahedron.
	The carbon atom can share electrons with other carbon atoms.	Two adjacent carbon atoms can share one, two, or three pairs of electrons.

Topics	Understandings and Fundamental Concepts	Supplementary Information
	The covalent bonding results in compounds that are molecular in structure.	
B. Structural formulas	The covalent bond is usually represented by a short line (or dash) representing one pair of shared electrons. A formula showing the bonding in this manner is known as a structural formula.	It is important to remember that molecules of organic compounds are three-dimensional in nature. *Teachers are urged to use models to represent the three-dimensional structure of organic compounds.*
C. Isomers	Compounds which have the same molecular formula but different structures are called isomers.	For example, the compounds CH_3CH_2CHO (propanal) and CH_3COCH_3 (acetone) are isomers, both having the molecular formula C_3H_6O. *(Students are expected to recognize isomerism in compounds.)* As the number of atoms in the molecule increases, the possibilities of more spatial arrangements (thus, the number of isomers) increase.
D. Saturated and unsaturated compounds	Organic compounds in which carbon atoms are bonded by the sharing of a single pair of electrons are said to be saturated compounds. Organic compounds containing two adjacent carbon atoms bonded by the sharing of more than one pair of electrons are said to be unsaturated compounds.	A bond formed between carbon atoms by the sharing of one pair of electrons is referred to as a single bond. A bond between carbon atoms by the sharing of two pairs of electrons is referred to as a double bond. A bond formed between carbon atoms by the sharing of three pairs of electrons is referred to as a triple bond.
III. Homologous Series of Hydrocarbons	The study of organic chemistry is simplified by the fact that organic compounds can be classified into groups having related structures and properties. Such groups are called homologous series. Compounds containing only carbon and hydrogen are known as hydrocarbons. Most carbon compounds are named from, and can be considered as related to, corresponding hydrocarbons.	Each member of an homologous series differs from the one before it by a common increment. As the members of a series increase in molecular size the boiling points and freezing points increase due to the increase in the van der Waals forces. (See Unit 3, Section III, C, p. 17.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
A. Alkanes	The series of saturated hydrocarbons having the general formula C_nH_{2n+2} is called the alkane series.	*Students should be able to recognize the names and formulas of the first five members of this series: methane, ethane, propane, butane, and pentane.* The alkane series is also called the methane series or the paraffin series. The alkane series show isomerism beginning with the fourth member (butane, C_4H_{10}). *In naming organic compounds the I.U.P.A.C. rules of nomenclature should be followed. The names of the isomers of butane and pentane will be subject to examination.*
B. Alkenes	The series of unsaturated hydrocarbons containing one double bond and having the general formula C_nH_{2n} is called the alkene series. In the I.U.P.A.C. (International Union of Pure and Applied Chemists) system of nomenclature the alkenes are named from the corresponding alkane by changing the ending "-ane" to "-ene."	*Students should be able to recognize the names and formulas of the first four members of this series: ethene, propene, butene, and pentene.* The alkene series is also called the ethylene series or the olefin series. There are series of hydrocarbons containing more than one double bond, such as the dienes. These, however, are not members of the alkene series.
C. Alkynes	The series of unsaturated hydrocarbons containing one triple bond and having the general formula C_nH_{2n-2} is called the alkyne series. In the I.U.P.A.C. system of nomenclature the alkynes are named from the corresponding alkane by changing the ending "-ane" to "-yne."	*Ethyne (acetylene) is the only member of this series that will be subject to examination.* The common name of the first member of this series, "acetylene," C_2H_2, is still in general use and should be familiar to students. The alkyne series is also called the acetylene series.
D. Benzene series	The benzene series is a series of aromatic hydrocarbons having the general formula C_nH_{2n-6}.	*Benzene and toluene are the only members of this series that will be subject to examination.* The simplest member of the benzene series is benzene, C_6H_6. The second member is toluene, C_7H_8, $(C_6H_5CH_3)$.

Topics

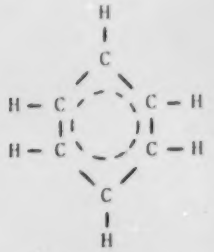
Understandings and Fundamental Concepts

Supplementary Information

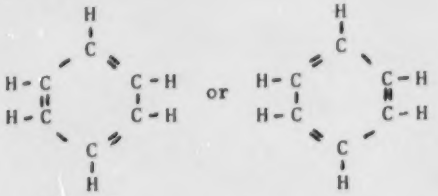
All of the carbon-carbon bonds in the benzene ring are the same, and they have structure and properties intermediate between single bonds and double bonds.

Benzene is rather unreactive and in many of its reactions behaves like a saturated hydrocarbon.

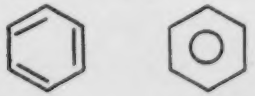
The structure of benzene is represented as:



This represents a "super position" or average of single and double bonds as shown below.



For simplicity, the chemist often uses either one of the structures shown below.



IV. Other Organic Compounds

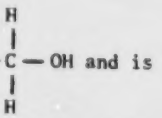
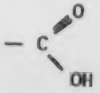
Other homologous series of organic compounds occur when one or more hydrogen atoms of a hydrocarbon have been replaced by other elements.

These compounds are usually named from their corresponding hydrocarbons, but are not necessarily prepared directly from the hydrocarbon itself.

A functional group is a particular arrangement of a few atoms which gives characteristic properties to an organic molecule.

Organic compounds can often be considered as being composed of one or more functional groups attached to a hydrocarbon group.

Students should be able to recognize primary alcohols and organic acids by their functional groups.

Topics	Understandings and Fundamental Concepts	Supplementary Information
A. Alcohols	In alcohols, one or more hydrogens of a hydrocarbon have been replaced by an -OH group.	No more than one -OH group can be attached to one carbon atom under ordinary conditions. The alcohols are not bases. The -OH group of an alcohol does not form a hydroxide ion in aqueous solution.
1. Primary alcohols	In primary alcohols, one -OH group is attached to the end carbon of a hydrocarbon.	Since the functional group can be the end group of any hydrocarbon, the typical alcohol is frequently represented as R-OH, where "R" represents the rest of the molecule. The end group of a primary alcohol has the structural formula:
	Primary alcohols contain the functional group $-CH_2OH$.	 and is frequently written as $-CH_2OH$.
	In the I.U.P.A.C. system of nomenclature, primary alcohols are named from the corresponding hydrocarbon by replacing the final "-e" with the ending "-ol."	*Students should know the names of the five primary alcohols: methanol, ethanol, propanol, butanol, and pentanol.* The common names of the alcohols were formerly derived from the name of the corresponding hydrocarbon by changing the ending "-ane" to "-yl" and adding the name "alcohol." Thus, CH_3OH, methanol, was called methyl alcohol. *It is recommended that the I.U.P.A.C. nomenclature be used.*
B. Organic acids	Organic acids contain the functional group - COOH.	Acids are represented by the general formula R-COOH, except for the first member. The structural formula of the acid group is: 
	In the I.U.P.A.C. system of nomenclature organic acids are named from the corresponding hydrocarbons by replacing the final "-e" with the ending "-oic" and adding the name "acid."	The first two members of this series, methanoic acid, HCOOH, and ethanoic acid, CH_3COOH, are more familiarly known by their common names, formic acid and acetic acid.

Topics	Understandings and Fundamental Concepts	Supplementary Information
		Students should know both the I.U.P.A.C. name and the common name of the first two members, methanoic (formic) acid and ethanoic (acetic) acid.
V. Organic Reactions	Organic reactions generally take place more slowly than inorganic reactions. Organic reactions frequently involve only the functional groups of the reacting species, leaving the greater part of the reacting molecules relatively unchanged during the course of the reaction.	*Minimum requirements for the study of different organic reactions are indicated in the following subsections of this section. It is expected that teachers will discuss the significance of these reactions.*
A. Substitution	Substitution means replacement of one kind of atom or group by another kind of atom or group. For saturated hydrocarbons, reactions (except for combustion and thermal decomposition) necessarily involve replacement of one or more hydrogen atoms.	The hydrogen atoms of saturated hydrocarbons can be replaced by active halogen family atoms. The general term for these reactions is halogen substitution and the products are called halogen derivatives. *In naming halogen derivatives of the hydrocarbons, the I.U.P.A.C. rules of nomenclature should be followed. Only halogen derivatives of the first five alkane hydrocarbons will be subject to examination.*
B. Addition	Addition usually involves adding one or more atoms at a double or triple bond, resulting in saturation of the compound. Addition is characteristic of unsaturated compounds. Because addition reactions take place more easily than substitution reactions, unsaturated compounds tend to be more reactive than saturated compounds.	Some addition reactions are about as fast as the reactions between ions. Alkynes are more reactive than alkenes. The addition of hydrogen to an unsaturated substance is called hydrogenation. This reaction usually requires the presence of a catalyst and a raised temperature. The addition of chlorine and bromine (usually iodine doesn't add) takes place at room temperature. The compounds formed are also called halogen derivatives.

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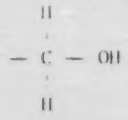
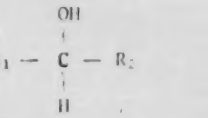
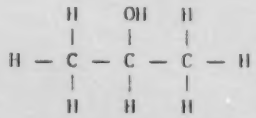
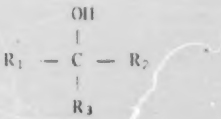
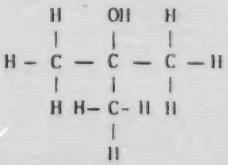
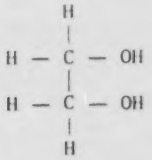
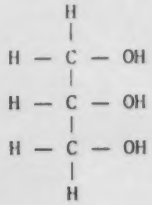
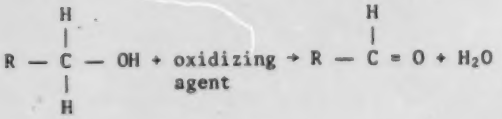
Topics	Understandings and Fundamental Concepts	Supplementary Information
C. Fermentation	In the fermentation process, enzymes produced by living organisms act as catalysts.	A common fermentation product, ethanol, results from the fermentation of sugar. For example: $C_6H_{12}O_6 \xrightarrow{\text{zymase (from yeast)}} 2C_2H_5OH + 2CO_2$
D. Esterification	Esterification is the reaction of an acid with an alcohol to give an ester and water. acid + alcohol = ester + water	Esterification is not an ionic reaction. Esters are covalent compounds. Esters usually have pleasant odors. The aromas of many fruits, flowers, and perfumes are due to esters.
	Esterification proceeds slowly and is reversible.	Fats are esters derived from glycerol and long-chain organic acids.
E. Saponification	The hydrolysis of fats by bases is called saponification.	To make soap, fat (a glycerol ester) is saponified by hot alkali. The products are soap (a salt of an organic acid) and glycerol.
F. Oxidation	Saturated hydrocarbons react readily with oxygen under conditions of combustion.	*Teachers may wish to discuss the combustion of hydrocarbon fuels.* In an excess of oxygen, hydrocarbons burn completely to form carbon dioxide and water. $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$ Burning in a limited supply of oxygen may produce carbon monoxide and carbon as well. $2\,CH_4 + 3O_2 \longrightarrow 2CO + 4H_2O$ $CH_4 + O_2 \longrightarrow C + 2H_2O$
G. Polymerization	Polymerization involves the formation of a large molecule from smaller molecules.	Synthetic rubbers, plastics such as polyethylene, and other large chain molecules synthesized by man are polymers. In nature polymerization occurs in the production of proteins, starches, and other chemicals by living organisms.

(Questions based on this material may appear only in Part II of the Regents examination.)

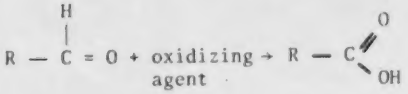
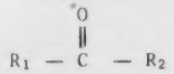
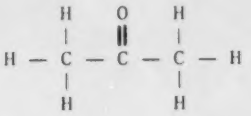
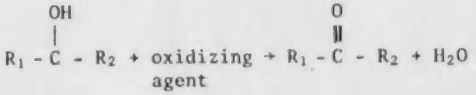
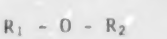
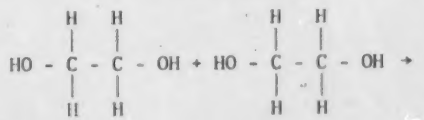
Teachers who elect this unit may wish to teach materials outlined here in the appropriate places in Unit 8.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Alcohols	Alcohols contain the functional group -OH. Alcohols can be classified according to the number of -OH groups in the molecule. They may also be classified according to the number of carbon chains attached to the carbon having the -OH group.	"R" represents a hydrocarbon group.
A. Monohydroxy alcohols	Monohydroxy alcohols have one -OH group.	
1. Primary alcohols	The primary alcohol group is  (See Unit 9, IVA, 1, p. 65.)	Methanol and ethanol are common primary alcohols.
2. Secondary alcohols	The secondary alcohol group is 	An example is 2-propanol (isopropanol). 
3. Tertiary alcohols	The tertiary alcohol group is 	An example is 2-methyl-2-propanol (tertiary butanol). 
B. Dihydroxy alcohols	Compounds containing two -OH groups are known as dihydroxy (dihydric) alcohols, or glycols.	The most important glycol is 1, 2-ethanediol, commonly called ethylene glycol. It has the structural formula:

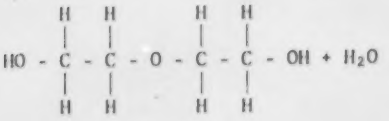
(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
		
C. Trihydroxy alcohols	Compounds containing three -OH groups are known as trihydroxy (trihydric) alcohols.	The most important trihydroxy alcohol is 1, 2, 3-propanetriol (glycerol), having the structural formula:  *Students should be able to recognize the name and formula of glycerol.* *Students who have studied biology will be familiar with glycerol and its relationship to fats.*
II. Aldehydes	Aldehydes contain the functional group In the I.U.P.A.C. system of nomenclature, aldehydes are named from the corresponding hydrocarbons by replacing the final "-e" with the ending "-al." Primary alcohols can be oxidized to aldehydes.	Aldehydes are represented by the general formula R-CHO where "R" is a hydrocarbon group except for the first member. The aldehyde of primary importance is methanal, HCHO, which is generally referred to by its common name, formaldehyde. 

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
	Aldehyde groups are easily oxidized to acids.	
III. Ketones	Ketones contain the functional group 	R_1 and R_2 are hydrocarbon groups. An important ketone, widely used as a solvent, is propanone,  generally referred to by its common name, acetone.
	Secondary alcohols can be oxidized to ketones.	
IV. Ethers	The functional group of an ether is 	R_1 and R_2 are hydrocarbon groups. Diethyl ether, $C_2H_5OC_2H_5$, is used as an anaesthetic.
	Primary alcohols can be dehydrated to give ethers.	
V. Polymers	A polymer is composed of many repeating units called monomers. Polymerization is the process of joining monomers. Polymers may be formed by condensation or by addition polymerization.	Starch, cellulose, and proteins are natural polymers. Nylon and polyethylene are synthetic polymers.
A. Condensation	Condensation polymerization results from the bonding of monomers by a dehydration reaction.	A condensation process may be illustrated as:  monomer monomer

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
		 dimer This process may be repeated to give a long-chain polymer. The prerequisite for this is that the starting material (monomer) has at least two functional groups. Silicones, polyester, polyamides, phenolic plastics, and nylons are all examples of condensation polymers. *Students are expected to recognize a condensation reaction when illustrated by equations having structural formulas.*
B. Addition	An addition polymerization results from the joining of monomers of unsaturated compounds by "opening" double or triple bonds in the carbon chain.	An addition process may be illustrated as: $nC_2H_4 \rightarrow (-C_2H_4-)_n$. Vinyl plastics such as polyethylene and polystyrene are examples of addition polymers.

(Questions based on this material may appear only in Part II of the Regents examination.)

Teachers who elect this unit may wish to teach materials outlined here in the appropriate places in the syllabus.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Chemical Theory and Industry	Pure research is directed toward the acquiring of knowledge without the consideration of the immediate practical application of that knowledge. The scientist's search for truth often leads to the benefit of mankind. Industrial processes are concerned with obtaining a maximum yield of products with maximum economic efficiency. An understanding and proper use of the principles involved in an industrial process can materially affect the quality, yield, and cost of the product.	
II. Industrial Applications		
A. Equilibrium and reaction rates	An application of the factors (such as concentration, temperature, pressure, and catalysts) affecting the rates of reaction, as well as equilibrium conditions, make many industrial processes practical.	
1. Haber process	Ammonia is produced by the reaction: $N_2 + 3H_2 = 2NH_3 + \text{heat}$. An increase in pressure and decrease in temperature will shift the equilibrium point in this reaction in the direction that favors the formation of more ammonia. Iron with traces of other materials is the catalyst used in the production of NH_3.	The rate of formation of ammonia is slow at low temperatures, so that the temperature must be raised to increase the rate of formation. However, since this is an exothermic reaction, high temperatures decrease the yield of ammonia. In practice, a compromise temperature of 400-600°C and a pressure of 200-1000 atmospheres are used to secure reasonable yields. The rate of formation of ammonia is increased by adding a trace of aluminum oxide or silicon oxide along with potassium oxide to the iron catalyst.
2. Contact process	Sulfur or sulfide ores such as iron pyrites are burned to produce SO_2. $S + O_2 \rightarrow SO_2 + \text{heat}$.	

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(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
	The sulfur dioxide is then oxidized to sulfur trioxide by the reaction $2SO_2 + O_2 = 2SO_3 + \text{heat}$.	Platinum or vanadium pentoxide is used as a catalyst in the oxidation of SO_2.
	An increase in pressure and a decrease in temperature favor the formation of SO_3.	The rate of formation of SO_3 is slow at low temperatures so the temperature must be raised to increase the rate of formation. However, since this is an exothermic reaction, high temperatures decrease the yield of SO_3. In practice, the temperatures are adjusted between 450-575°C to attain reasonable yields of SO_3.
	The sulfur trioxide formed is then absorbed by sulfuric acid. $H_2SO_4 + SO_3 \rightarrow H_2SO_4 \cdot SO_3$ or $H_2S_2O_7$	The SO_3 is absorbed very slowly in water but very readily in concentrated sulfuric acid.
	$H_2S_2O_7$ is then diluted with water to produce sulfuric acid. $H_2S_2O_7 + H_2O \rightarrow 2H_2SO_4$	
B. Redox	Principles involving the "competition" for electrons by atoms help to explain many processes used by industry. Many basic redox reactions are employed to reduce or oxidize elements from compounds. Redox reactions are also used in refining and protecting metals.	
1. Reduction of metals	The ores of most metals contain the metal in an oxidized state. The form in which a metal occurs is related to the chemical activity of the metal and the stability of its compounds.	In general, compounds of metals that occur in nature have high stability and low solubility in water.
	Reduction of the ore is necessary to obtain the metal, the method of reduction depending on the activity of the metal and the type of ore.	
	The most active metals are obtained from their fused compounds by electrolytic reduction. $2NaCl\ (\text{fused}) + \text{electricity} \rightarrow 2Na + Cl_2$	Groups IA and IIA metals are obtained by electrolysis of their fused salts.

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
	Metals that form relatively stable compounds can be liberated from their compounds by stronger reducing agents. For example, in the production of chromium, aluminum reduces the chromium oxide.	
	$2\ Al + Cr_2O_3 \rightarrow Al_2O_3 + 2Cr$	
	Moderately active metals like zinc and iron are obtained by reduction of their oxides by carbon (coke) or carbon monoxide.	Many metals are found as sulfides and carbonates. These ores are converted to the oxide form and then reduced by carbon.
	$ZnO + C + heat \rightarrow Zn + CO$	
	$Fe_2O_3 + 3CO + heat \rightarrow 3CO_2 + 2Fe$	
2. Corrosion	Corrosion is a gradual attack on a metal by its surroundings with the result that the usefulness of the metal may be destroyed.	
	Corrosion is a redox reaction. Moisture, some gases in the air, and some chemicals contribute to corrosion.	
	Metals like aluminum and zinc form self-protective coatings.	Aluminum is more susceptible to corrosion than iron. However, the corrosion of aluminum is not a serious problem. The aluminum oxide formed can adhere tightly to the uncorroded aluminum beneath it and provide a protective layer that prevents further corrosion.
		In the case of iron, the oxide formed from corrosion has no tendency to adhere to the metal. It constantly flakes off, exposing fresh iron to corrosion.
	Metals that corrode easily, like iron, may be protected by a variety of methods.	Some methods include plating with self-protective metals like aluminum and zinc, or plating with corrosion resistant metals like chromium and nickel. In some instances, a more active metal like magnesium corrodes preferentially when it is connected to iron. The alloying of iron with corrosion resistant metals like nickel and chromium has produced stainless steel. Coating iron with paints, oils, or glass (porcelain) has proven effective against corrosion.

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
3. Batteries	Batteries are chemical cells in which spontaneous redox reactions are used to provide a source of electrical energy.	
a. Lead-acid battery	The change in oxidation states of lead is utilized to produce electric energy.	When fully charged, the positive and negative electrodes of this widely used battery are PbO_2 and Pb respectively. The electrolyte is a sulfuric acid solution. The overall reaction is $Pb + PBO_2 + 2H_2SO_4 \underset{\text{charge}}{\overset{\text{discharge}}{\rightleftharpoons}} 2PbSO_4 + 2H_2O$
	Reversal of the chemical process recharges the storage battery for further use.	In the battery, sulfuric acid is a reactant and therefore the concentration of this electrolyte decreases as the battery discharges.
b. Nickel-cadmium battery	The changes in oxidation states of nickel and cadmium are used to produce a source of electric energy.	The positive and negative electrodes of this compact rechargeable battery are $Ni(OH)_3$ and Cd, respectively. The electrolyte is a solution of KOH. The probable reaction is $2Ni(OH)_3 + Cd \underset{\text{charge}}{\overset{\text{discharge}}{\rightleftharpoons}} 2Ni(OH)_2 + Cd(OH)_2$
		In this battery, the electrolyte concentration remains the same.
C. Petroleum	Petroleum is a complex natural mixture of many hydrocarbons. Many petroleum products are used as fuels.	Petroleum is the important starting material for many chemical products like plastics, textiles, rubber, and detergents. Natural gas (mostly methane), a common fuel, is often found with petroleum. Common bottled gases like propane and butane are obtained from petroleum.

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Fractional distillation	Because hydrocarbons differ in boiling points, petroleum can be separated into fractions of simpler mixtures by distillation.	The uses of these different fractions are many and varied. The important fractions are gasoline, kerosene, fuel oil, lubricating oils and greases, paraffin wax, and asphalt.
2. Cracking	In cracking, some large molecules of hydrocarbons (higher boiling fractions) are broken down into smaller molecules of lower boiling points.	The cracking process is utilized to increase the yield of valuable fractions. Two examples are gasoline and fuel oil. Generally, cracking requires the use of a catalyst (example: oxides of aluminum and silicon) and increased temperature.

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(Questions based on this material may appear only in Part II of the Regents examination.)

Teachers who elect this unit may wish to teach materials outlined here in the appropriate places in Unit 2.

Topics	Understandings and Fundamental Concepts	Supplementary Information
I. Artificial Radioactivity	Elements can be made radioactive by bombarding their nuclei with high energy particles such as protons, neutrons, and alpha particles.	When the element beryllium, for example, is bombarded by protons, the nuclear reaction that takes place is $^{9}_{4}Be + {}^{1}_{1}H \rightarrow {}^{6}_{3}Li + {}^{4}_{2}He.$
A. Artificial transmutation	Bombardment of nuclei by accelerated particles cause nuclei to become unstable and may result in the formation of isotopes, or new elements.	Natural radioactivity is an example of natural transmutation. *Students should recognize transmutation when illustrated by an equation:* $^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{15}P + {}^{1}_{0}n.$
1. Accelerators	Electric and magnetic fields are used to accelerate charged particles.	Accelerators are used to give charged particles sufficient kinetic energy to overcome electrostatic forces and penetrate the nucleus. *Knowledge of details of construction and operation of accelerators is not required.*
I. Nuclear Energy	In nuclear reactions mass is converted to energy.	Nuclear reactions involve energies much greater than ordinary chemical reactions. The energy changes may be a million or more times larger.
A. Fission reaction	Fission is brought about by neutron capture and this results in fission fragments, liberation of energy, and release of two or more neutrons. The liberation of energy is the result of conversion of mass into energy.	Only elements of high atomic numbers can be fissioned. When a heavy element fissions, the new elements formed have a more stable configuration because of the greater binding energy per nucleon. In a nuclear reactor, the chain reaction can be controlled. In an atomic bomb, the chain reaction is not controlled.

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Fuels	Uranium-233, uranium-235, and plutonium-239 are fissionable.	Natural uranium (99.3% uranium-238 and 0.7% uranium-235) and enriched uranium (3 to 4 percent enrichment with uranium-235) are commonly used as fuels.
		Uranium-233 produced from thorium-232 and plutonium-239 produced from uranium-238 are obtained as fuels in breeder reactors. Breeder reactors produce more fuel than is consumed.
2. Moderators	For efficient nuclear fission, it is necessary to slow down the speed of the neutrons. Moderators are materials that have the ability to slow down neutrons quickly with little tendency to absorb them.	Particles of similar mass such as hydrogen and its isotope deuterium have been found effective as moderators.
	The neutrons are slowed down most effectively by a head-on collision with a particle of similar mass.	Water, heavy water, beryllium, and graphite are commonly used as moderators.
3. Control rods	The fission process in a reactor can be controlled by adjusting the number of neutrons available.	Boron and cadmium are commonly used in control rods because they absorb neutrons very well.
4. Coolants	Coolants are used to keep the temperatures generated by fission at reasonable levels within the reactor and to carry heat to heat exchangers and turbines.	Water, heavy water, air, helium, carbon dioxide, molten sodium and molten lithium are examples of coolants.
		In some reactors the coolant also serves as the moderator.
5. Shielding	The internal shield protects the walls of the reactor from radiation damage. The external shield serves to protect the personnel from radiation.	Steel lining is used for internal shielding and high density concrete for external shielding.
B. Fusion reaction	Fusion is the process of combining two light nuclei to form a heavier one. The energy released in a fusion reaction is much greater than in a fission reaction.	When two light nuclei fuse into a heavier nucleus at high temperatures and pressures, an element of more stable configuration (with greater binding energy per nucleon) is formed. The mass of the heavier nucleus formed is less than the sum of the masses of the lighter nuclei; the difference in mass is converted into energy.
		Solar energy is probably the result of the fusion of ordinary hydrogen atoms to form helium. The nuclear reaction of a hydrogen bomb utilizes fission as a trigger for fusion.

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
1. Fuels	The isotopes of hydrogen, $^{2}_{1}H$ (deuterium) and $^{3}_{1}H$ (tritium), are used as fuels.	Heavy water (deuterium oxide) is obtained by concentrating the trace quantities present in water. Tritium is made by the nuclear reaction $^{6}_{3}Li + {}^{1}_{0}n \rightarrow {}^{3}_{1}H + {}^{4}_{2}He.$
2. High energy requirement	Since each nucleus carries a positive charge, nuclei repel one another with increasing strength as they are moved closer together. Consequently, for the nuclei to interact, they must have enough kinetic energy to overcome the repulsion.	Since the magnitude of repulsion increases with charge, only the nuclei of lowest possible charge can be used. Fusion with ordinary hydrogen, however, is very slow. Fusion reactions involving deuterium, or deuterium and tritium, are useful sources for the release of energy. Of these, the most rapid reaction is between deuterium and tritium.
	The thermonuclear approach, through the use of very high temperatures, appears to be very promising for controlled fusion.	Technical problems with nuclear fusion, such as the requirement of extremely high temperatures and their containment, continue to challenge nuclear scientists and engineers.
C. Radioactive wastes	Fission products from nuclear reactors are intensely radioactive and cannot be discarded. They must be stored for a long time or disposed of in special ways.	Solid and liquid wastes such as strontium-90 and cesium-137 are encased in special containers for permanent storage underground or in isolated areas. Low level radioactive wastes may be diluted and released directly into the environment. Gaseous radioactive wastes such as radon-222, krypton-85, and nitrogen-16 are stored for decay to safe levels and then dispersed into the air.
D. Uses of radioisotopes		
1. Based on chemical reactivity	Since radioisotopes are chemically similar to stable isotopes of the same element, they can be used as tracers to follow the course of a reaction without seriously altering the chemical conditions.	Many organic reaction mechanisms are studied by the use of carbon-14 as a tracer.

(Questions based on this material may appear only in Part II of the Regents examination.)

Topics	Understandings and Fundamental Concepts	Supplementary Information
2. Based on radioactivity	Radioisotopes are used in medical diagnosis, therapy, food preservation, and as a means of measuring the physical dimensions of many industrial products.	Isotopes with very short half-lives and which will be quickly eliminated from the body are used for diagnostic injections. Technetium-99 is used for pinpointing brain tumors. Iodine-131 is used for diagnosing thyroid disorders. Radium and cobalt-60 are used in cancer therapy. Radiation kills bacteria, yeasts, molds, and insect eggs in foods, permitting the food to be stored for a much longer time.
3. Based on half-life	Radioisotopes give a fairly consistent method of dating some geologic events.	The ratio of uranium-238 to lead-206 in a mineral can be used to determine the age of the mineral. The ratio of carbon-14 in a sample, to that in the atmosphere, permits the age of the sample to be determined.

(Questions based on this material may appear only in Part I of the Regents examination.)*

THIS UNIT IS NOT INTENDED TO INCLUDE ALL OF THE REGENTS CHEMISTRY LABORATORY REQUIREMENTS. SEE INTRODUCTION P. ii.

It is not intended that the materials outlined here limit the teacher in the laboratory. Rather, the topics outlined here represent the minimum requirements for testing if this unit is elected in Part II of the Regents Chemistry Examinations.

The skills and activities in this unit should be developed and reinforced throughout the year by the students' laboratory experiences.

STUDENTS SHOULD BE FAMILIAR WITH GENERAL LABORATORY SAFETY PROCEDURES.

Activities	Minimum Requirements	Supplementary Information
I. Measurement	Students should be able to use common measuring devices: balance, graduated cylinder, thermometer, buret, etc. Students should be able to round off numbers according to the correct number of significant figures.	If possible, students should be able to interpolate to 1/10th of the smallest scale division of a measuring device. The accuracy of a measurement or calculated result can be indicated by the use of significant figures. When adding or subtracting, the answer should be rounded off to contain the least accurately known figure as the final one. Examples: Add: 32.6 + 431.33 + 6144.212 = 6608.142 = 6608.1 Subtract: 531.46 - 86.3 = 445.16 = 445.2 When multiplying or dividing, the answer should be rounded off to contain only as many significant figures as are contained in the least accurate number. Examples: Multiply: 1.36 x 4.2 = 5.712 = 5.7 Divide: $\frac{5.1}{2.13}$ = 2.39 = 2.4

(Questions based on this material may appear only in Part II of the Regents examination.)

Activities	Minimum Requirements	Supplementary Information
		When the number dropped is 5 or more, the preceding number is increased by 1. For example, 2.4179 to three significant figures becomes 2.42.
	Students should be able to determine percent error wherever applicable.	Percent error $= \frac{\text{observed value - accepted value}}{\text{accepted value}} \times 100$ Observed value is the experimentally measured value or the value calculated from the experimental results. The "accepted value" is the most probable value taken from generally accepted references.
II. Laboratory Skills	Students should acquire at least the following basic laboratory skills:	
	• The proper adjustment and use of the burner or heat source	
	• How to cut, bend, and firepolish glass tubing	
	• The accepted use of a funnel and filter paper to separate a mixture	
	• The safe technique for dilution of acids and other reagents	
	• The correct technique of removing solid chemicals from containers	
	• The safe method of pouring liquids	
	• Safely heating materials in test tubes, beakers, and flasks	
	• The evaporation of a liquid or the drying of a solid	
	• Identification of common laboratory apparatus	Some examples are: Test tube, beaker, Erlenmeyer flask, evaporating dish, graduated cylinder, volumetric flask, buret, pipette, medicine dropper, bunsen burner, wire gauze, pipe stem triangle, crucible tongs, test tube holder, ring, ring stand, funnel and watch glass.

(Questions based on this material may appear only in Part II of the Regents examination.)

Activities	Minimum Requirements	Supplementary Information
Laboratory Activities	Students should do a wide variety of activities in the laboratory. They should be able to:	*Teachers may draw upon personal preferences and the availability of apparatus to perform these activities. The materials outlined here are suggestions.*
	• Carry out an activity involving phase change and interpret a simple heating or cooling curve based on this activity.	Ice or paradichlorobenzene or naphthalene may be used to determine some portions of a heating or cooling curve.
	• Measure heats of simple reactions in calories	A simple activity in calorimetry may be done using water as a calorimetric liquid and a burning candle to determine the heat of combustion of the candle.
	• Identify endothermic and exothermic processes	It is possible to measure the temperature changes associated with certain chemical reactions or solution processes. The dissolution of ammonium chloride is endothermic and that of sodium hydroxide is exothermic.
	• Do a solubility experiment and draw a solubility curve	Salts such as sodium nitrate, potassium nitrate, and potassium chlorate show large increases in solubility as the temperature increases.
	• Identify differences between inorganic and organic substances	Differences in solubility, melting point, stability, and electrical conductivity may be qualitatively demonstrated using sodium chloride and sucrose.
	• Identify metallic ions from a flame test	Salts of lithium, sodium, potassium, calcium, strontium, and copper may be used.
	• Find the percent by mass of water in a hydrated crystal.	The hydrate, $BaCl_2 \cdot 2H_2O$, gives excellent results on heating to constant mass in a crucible.

(Questions based on this material may appear only in Part II of the Regents examination.)

Activities	Minimum Requirements	Supplementary Information
	• Determine the volume of a mole of gas in a reaction.	The hydrogen gas generated from the reaction of approximately 0.05g magnesium with excess dilute HCl may be collected in a 50 ml buret or a eudiometer tube. The volume of the hydrogen may then be corrected to STP and related to the moles of magnesium used in the reaction.
	• Titrate a dilute solution of an acid (or base) of unknown concentration against a dilute solution of a base (or acid) of known concentration.	The end point of the titration may be determined by use of an appropriate indicator such as methyl orange. Dilute solutions of sodium hydroxide or potassium hydroxoide may be titrated against dilute hydrochloric or sulfuric acids. If burets are not available, calibrated medicine droppers or pipettes can be used.
IV. Laboratory Reports	Students should be able to write concise reports of laboratory experiences. They should be able to: • Organize information in a logical manner • Put appropriate data in tables and graphs whenever possible • Make a list of observations • Draw conclusions based on observations	